Chemistry 101: The Complete Notes

Dr. Patrick Mills, Joliet Junior College
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See the Course Website (http://www3.jjc.edu/staff/pmills) for specific test dates and other important information

Legend

You will often find specific icons embedded within the notes. These respective symbols alert the student to the following:

- **Key** represents a key fact or other piece of information, such as the definitions of an element and a compound.

- **Magician Hat** represents a useful trick the student will likely find useful, such as an 'EZ' way to convert between grams and moles for a substance.

- **Professor** alerts the student to an important relationship between micro and macro scale properties or phenomena with respect to the material under discussion.

- **Star** such material provides a link to interesting (briefly discussed) supplemental material, often beyond the scope of the course syllabus.

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Cover Art: The ‘Fundamental Sun’ (Atom) STM image
Why Chemistry?

“What’s my motivation?”

Why are you sitting in this class? In other words, why chemistry?

Task: Write down as many reasons as you can that explain why you are taking this class:
(We will also justify responses relating to 3rd party requirements during the session)

Professional programs that benefit directly from a background in chemistry

1. Nursing and allied health (pre-pharm., pre-med., pre-dentistry)

Example: Chlorothiazide (Diuril) is ordered b.i.d. for an infant weighing 6.5 kg. It is supplied in elixir form 100 mg/tsp. The recommended dosage for Diuril is 25 mg/kg/day. How many cc’s should the nurse give to the child for each dose?

   A. 6.15 cc.
   B. 8.13 cc.
   C. 4.06 cc.
   D. 0.81 cc.
2. Engineering (mechanical, civil, chemical, electrical)

Example: Your company decides to import child safety seats manufactured in Asia. Unfortunately, the safety guidelines for the seats are quoted in ‘metric’ units. The label reads: “Do not exceed a 150 N load” and you must use this information to determine the maximum weight a child must not exceed in order to be protected during a collision at 55 mph. Can you do it? A child’s life, not to mention the financial future of your employer, may depend on your ability to solve questions such as this.

3. Everyday / Real life situations

Example: It is time to re-carpet your 12 ft x 24 ft. family room. You visit a few carpet stores and select a brand that costs $20.50 per square meter. The sales person quotes you a total price of $749 – is this price fair, or have you just been taken advantage of?

** We will return to and solve each of these three problems at some point during the course

Discussion:

What do all three of the above examples have in common?

Which professions (or professionals) utilize such skills most commonly?

Hint, “I pretend to be one on occasion”

“We’re the first link”
The “Cognitive Elite”

Discussion: What do you think the phrase “cognitive elite” actually means?

Data from ‘The Emergence of a Cognitive Elite’ (Chapters 1 and 2 of *The Bell Curve*).

- People with IQ’s of > 120 (the top 10%) preferentially enter the 10 or so ‘High IQ professions’ discussed above.

- Developing good cognitive skills is essential to entering and being successful within the ‘High IQ’ and related professions. **We are the first link in the chain**

Example: Medical careers

\[ 	ext{Increased Problem Solving Ability} \]

Take home message: People with good cognitive / problem solving skills preferentially find employment within fields of their choosing that are financially rewarding and/or intellectually satisfying.

A question of some importance: How can one’s cognitive skills be improved?

Answer(s):
The Role of Chemistry as a Prerequisite Course

Key facts and results:

**Fact:** The problem solving skills routinely utilized in the ‘high IQ’ and related professions (such as nursing, business management, accounting, etc.) are introduced, learnt and mastered during physical science courses.

**Result:** Professional programs and subsequent employers insist that their candidates have a background in one of the physical sciences – both for specific (allied health, engineering) and general (your family room carpet) reasons.

**Fact:** Study within any of the ‘high IQ fields’ will increase cognitive skills, but only the physical sciences do so via the study of fundamental, everyday phenomena so are of broad relevance and interest (we all interact with and benefit from the manipulation of matter on a daily basis after all).

**Result:** Chemistry (and physics) may be considered to be the ‘gatekeepers’ of cognitive learning – chemistry in particular introduces, develops and subsequently equips students with cognitive skills necessary to succeed in their chosen careers.

Take home message: While the direct relevance of chemistry to your chosen course of study may at times seem tenuous, remember that the cognitive skills developed during such programs of study are of significant importance to your professional development and employability. In essence, this is why you are here.
How Chemistry is Perceived & Skills Needed to Succeed in Chemistry

How Chemistry is Perceived:

Discussion: How did your friends and family respond when you told them you were taking a chemistry course this semester?

[“Frank” slide]

Study Skills Needed to Succeed in Chemistry:

Fact: As discussed above, chemistry is all about the student developing and learning to apply problem solving skills - your study habits should reflect this. Do NOT fall in to the trap of believing you can learn chemistry simply by memorizing the information from your text – you must practice applying this information, not just be familiar with it.

Result: Successful chemistry students typically spend most of their independent study time working assigned problems, not just reading about them. To learn chemistry you must do chemistry is a truism worth remembering. An analogy would be this: you read all the books out there on the subject of golf, but don’t get round to swinging a club – what do you think happens when you tee off for the first time?

Fact: Chemistry relies on a cumulative method of learning, i.e. theories learnt from week 1 onwards will be repeatedly applied all the way through the course. Thus, it is important that the student does not let any ‘gaps’ in their knowledge develop. This fact exemplifies the differences in philosophy between the sciences and arts, as art courses are often more modular in nature. Example: I overhead a student tell another: “Yeah, I blew off reading the first book in my English class, but read the second one and got a ‘B’”. This method of study is not recommended in chemistry!
Analogy: Building a tower

Result: Successful chemistry students typically have exemplary attendance records. In some cases they may not be the ‘best’ students, but guarantee themselves a better grade than more capable students, who in turn typically may miss as few as one or two lecture sessions (this is especially true with regard to 3 hr. class sessions).

Pictorial analogy of attendance vs cumulative knowledge

‘I missed a lab’     ‘I missed a lecture’     ‘I missed a couple of lectures’

Don’t ‘Swiss cheese’ or ‘torpedo’ your chances of passing the course because of missed work!

Take home message: Simply by attending class regularly and completing the HWK assignments you essentially guarantee yourself a passing grade for the course, while, due to the nature of the material, deviating from this approach may ensure the opposite
What is chemistry? What do Chemists do?

Task: In your own words describe what you consider chemistry to be, plus make a list of what you think the job of a chemist is:

What is chemistry?

“Official” definition of what chemistry is:
Key words:

**Matter:** “Stuff” – **anything with mass and volume.** Can you think of anything without mass or volume?

What are the basic building blocks of all matter, be it a diamond, a tree or the air around us?

‘High Tech’ science (STM or AFM, top left) is often based on simple ideas (gramophone, top right). Click logo for ‘flyby’.

**More recent atomic (STM) images**
Example: What is water made up from? How do you know?

Summary:

Atoms and molecules are MICROSCOPIC particles (they are very small)

A drop of water is a MACROSCOPIC particle (because you can see it, hold it in your hand etc.)
What do chemists do? In other words, what is the basic most, fundamental goal of every chemical investigation? **Hint:** Think how chemists express their findings..

“Official” definition of what chemists do:
Chemistry in action: Explaining what happens on your BBQ grill.

The burning of a charcoal brick on your backyard grill (MACRO) explained in terms of a balanced chemical equation (MICRO)

**Key**

ANY large (MACRO) scale chemical process can be described using a MICRO scale chemical equation featuring individual atoms and/or molecules

A cartoon representation of the reaction of the pertinent atoms and molecules; along with the Chemists’ description – a balanced chemical equation illustrating a single microscopic event.

**Cartoon**

**Balanced Equation**

This process is repeated many billions of times (MICRO) for the burning of a charcoal briquette (MACRO)
The Components of Matter

Review: What is matter?

Recall: “Chemistry is the study of matter and its properties, the changes matter undergoes and the energy associated with those changes”

Recap: There are 3 stable states of matter – solid (s), liquid (l) and gas (g).
Specific macro- and microscopic physical properties define the three* states of matter

<table>
<thead>
<tr>
<th>State of Matter</th>
<th>Macroscopic Description (observation)</th>
<th>Microscopic Description (chemical model)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Liquid</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Gas</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The state matter is in depends on the strength of the forces (chemical bonds) between the individual microscopic particles within the matter

Task: Rank the intermolecular forces present in steam, ice and water in order of increasing strength. Use the included figures as a guide.

Ranking
Changing between the 3 states of matter

Describe the relationship between the mpt. and bpt. of matter, with regard to *microscopic* processes, occurring at these specific temperatures

Example: The boiling of water to make steam ( $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$)

Water molecules change from liquid to gaseous state: physical change.
Physical and Chemical Properties – what’s the difference?

**Analogy**: We all possess ‘as is’ physical properties, or characteristics, that define us. For example, Dr. Mills is 5’11” and has green eyes.

As with people, each chemical also possesses a unique set of ‘as is’ **physical properties** that define it. For example, water is a clear, colorless, tasteless molecular material that has a fpt. of 0°C and a bpt. of 100°C.

**Chemical Properties**, in contrast, are a function of change (usually associated with a chemical reaction). For example, Iron (Fe) reacts with oxygen gas to form rust:

\[
4 \text{Fe (s)} + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3 (s)
\]

**Task**: Identify the flowing as either chemical or physical properties

<table>
<thead>
<tr>
<th>Property</th>
<th>Chemical or Physical</th>
</tr>
</thead>
<tbody>
<tr>
<td>Diamond is the hardest known substance.</td>
<td>Chemical or Physical</td>
</tr>
<tr>
<td>Charcoal burns to make CO(_2) (g)</td>
<td>Chemical or Physical</td>
</tr>
<tr>
<td>The statue of liberty turned ‘green’</td>
<td>Chemical or Physical</td>
</tr>
<tr>
<td>Copper is a good conductor of electricity</td>
<td>Chemical or Physical</td>
</tr>
<tr>
<td>Sugar dissolves in water*</td>
<td>Chemical or Physical</td>
</tr>
<tr>
<td>Melting of ice*</td>
<td>Chemical or Physical</td>
</tr>
</tbody>
</table>

Think up two more chemical properties of your own
**Elements and Compounds** – the further classification of *pure* matter

Task: State which of the following are *elements*, and which are *compounds*. When done, try to come up with a definition of what elements and compounds are.

<table>
<thead>
<tr>
<th>Material</th>
<th>Chemical Formula</th>
<th>Element or Compound?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>H₂O (l)</td>
<td></td>
</tr>
<tr>
<td>Oxygen gas</td>
<td>O₂ (g)</td>
<td></td>
</tr>
<tr>
<td>Pure silver coin</td>
<td>Ag (s)</td>
<td></td>
</tr>
<tr>
<td>Sugar crystals</td>
<td>C₆H₁₂O₆ (s)</td>
<td></td>
</tr>
<tr>
<td>Carbon dioxide gas</td>
<td>CO₂ (g)</td>
<td></td>
</tr>
</tbody>
</table>

**Elements:**

**Compounds:**
Compounds and elements can have either ‘giant’ or molecular structures:

‘Giant’: Repeating lattice of particles – usually strongly bound (high mpt.) solids.

Examples: sand (SiO₂), diamond (C), table salt (NaCl)

Molecular: a collection of independent molecular units (molecules will be discussed in more detail later). Usually (low mpt) liquids or gasses at room temp.

Definition: Molecule – a small, independent particle of matter made up from 2 or more atoms

Examples: water (H₂O), carbon dioxide (CO₂), Nitrogen gas (N₂)

Think of molecules like cars on the expressway – each car (molecule) is a separate, independent unit that contains a number of passengers (atoms). The cars (molecules) are free to move while the people (atoms) stay fixed inside.

‘Giant’ materials are like people (atoms) ‘locked’ in place at a very crowded concert, the DMV waiting room etc……
Review: A microscopic scale view of several materials is presented below. Label each using *elemental or compound* and *molecular* or ‘giant’ tags.

<table>
<thead>
<tr>
<th>Material</th>
<th>Structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water (H₂O (l))</td>
<td><img src="image" alt="Water Molecule" /></td>
</tr>
<tr>
<td>Steam (H₂O (g))</td>
<td><img src="image" alt="Steam Molecule" /></td>
</tr>
<tr>
<td>Silicon (Si (s))</td>
<td><img src="image" alt="Silicon Molecule" /></td>
</tr>
<tr>
<td>Sodium Chloride (NaCl)</td>
<td><img src="image" alt="Sodium Chloride Crystal" /></td>
</tr>
</tbody>
</table>

Details: Ice is a solid (crystalline) form of water (a molecular compound). How would you describe the structure of ice? Can you think of other similar examples?

More Details: Allotropes of an Element

**Example:** Carbon

- C\(_{(\text{diamond})}\)
- C\(_{(\text{graphite})}\)
- C\(_{60}\)
Pure Matter v Mixtures

Recap: **Pure matter** is classed as *either* an ELEMENT *or* a COMPOUND.

**Elements can have *either* Molecular or ‘giant’ structures.**
*Examples: N\(_2\) (g) (Nitrogen gas, molecular), Pb(s) (metallic lead, a ‘giant’ structure)*

**Compounds can also have *either* Molecular or ‘giant’ structures.**
*Examples: H\(_2\)O(l) (water, molecular), Fe\(_2\)O\(_3\)(s) (‘rust’ (iron oxide), a ‘giant’ structure)*

**Recall:** A molecule is an independent unit containing two or more atoms. Remember the car / passenger analogy. Molecules can exist as *either* elements *or* compounds

Mixtures

**ANY combination of different types of pure matter *‘placed together’* is defined as a mixture (eg. air, milk, pepsi).**

**Mixtures are NOT pure materials.** eg. Pure gold (Au) vs ‘white’ gold (Au+ Ag), or water (H\(_2\)O) vs pepsi (H\(_2\)O + sugar….)

Discussion: Air contains a number of different components – what are they? How would you describe what air is made up from using words like element, compound, gas, molecular etc.?
Task: Assign generic labels that describe to microscopic scale matter shown on the slide (e.g. ‘gaseous atomic element’ etc.)

Mixture Types

As viewed from a macroscopic perspective, mixtures are classified as either HOMOGENEOUS or HETEROGENEOUS

HOMOGENEOUS MIXTURES:

Examples:

HETEROGENEOUS MIXTURES:

Examples:
Discussion: Can you think of something that is both a homogeneous mixture and a solid?

Examples of Alloys:

Classification of Matter Flowchart
(Dr. Mills really likes this slide – why? Hint: Recall the fundamental job of a chemist)
Task: Use the ‘Classification of Matter’ flowchart (above) to classify the following:

1. The compressed gasses in a deep sea diver’s gas bottle (He(g) and O₂(g))

2. A ham and cheese omelet

3. An ice cube (made from pure water)

4. A ruby (Al₂O₃(s) with Cr³⁺ impurities)

Extra Credit: Ask me about the separation of mixtures assignment (based on background reading)
State whether the following are classified as elements, compounds or mixtures**:

**Diamond:**

**Carbon dioxide gas:**

**Air:**

**A cup of coffee:**

**Water:**

**Sand (SiO₂):**

**Oxygen gas:**

**include additional details for extra credit!**
Units of Measurement

Reading: Ch. 1 sections 6 & 7  Homework: Chapter 1: 51, 55*, 57*, 59, 61, 65, 67, 69*, 71*, 75, 77*, 81, 83*

* = ‘important’ homework question

Common Units

Discussion: List some common units of measurement we use on a daily basis. How did these units originate?

<table>
<thead>
<tr>
<th>Quantity measured</th>
<th>Familiar Unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
<td></td>
</tr>
</tbody>
</table>

Question: What are the ‘metric’ (S.I.) versions of the everyday units listed above?

<table>
<thead>
<tr>
<th>Quantity measured</th>
<th>Fundamental S.I. Unit (base unit)</th>
<th>Symbol</th>
</tr>
</thead>
</table>

Notes: SI base units are used to determine derived S.I. units, as discussed below. Some S.I. base units feature a decimal prefix – which one(s)?
Derived S.I. Units

Insert appropriate S.I. base units into an equation that defines the respective derived S.I. unit desired.

Example:

\[ \text{Area} = \text{length} \times \text{length} = \text{m} \times \text{m} = \text{m}^2 \]

\[ \Rightarrow \text{the derived S.I. unit for area is} \ \text{m}^2 \]

Determine derived S.I. units for the following quantities

<table>
<thead>
<tr>
<th>Quantity measured</th>
<th>Math involving S.I. base units</th>
<th>Derived S.I. unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Velocity (speed)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Density</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Force*</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Energy*</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*These are harder examples. To solve them start by inserting appropriate S.I. base units into an equation that defines the quantity sought.

Discussion: Why do scientists prefer the S.I. system?
Questions:

Is the S.I. unit of volume (m$^3$) reasonable for everyday applications? Why?

What unit of volume do chemists prefer? Why?

1 dm$^3$ = 1 L

More detail on the chemist’s volume unit
Scientific Notation and S.I. Prefixes

Large Quantities

Fact: Chemical problem solving *most often* involves using either very large or very small numbers (e.g. counting the number of molecules in a drop of water, or quoting the mass of the water drop in kilograms)

Recall: How many individual H$_2$O (l) molecules are there in a drop of water. Write this amount as a regular number:

Number H$_2$O (l) molecules in 1 drop water = ____________________________

Problem: How do we represent and manipulate such quantities in an ‘easier’ way?

Answer:

Overview Example: Consider the statement “*eight million people live in London*”. How can this quantity be *best* expressed numerically?

‘Everyday’:

‘Better’:
Examples: Write the following quantities using *regular numbers* and *powers of 10 (scientific notation)*. Try to do this without a calculator at first, then see the below tip for how to do this with your calculator’s [SCI] button.

**Just move the decimal point to the left until you get a single digit with decimals.** The *power of ten* is the number of places the decimal point moved. Example:

\[3000 = 3 \times 10^{\text{(number decimal places to left moved)}} = 3 \times 10^3\]

<table>
<thead>
<tr>
<th>Quantity</th>
<th>‘Regular’ quantity</th>
<th>‘Power of ten’ quantity (SCI)</th>
</tr>
</thead>
<tbody>
<tr>
<td>One hundred miles</td>
<td></td>
<td></td>
</tr>
<tr>
<td>One thousand students</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Five million people</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Twenty million dollars</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Five and a half billion people</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
TIP: Scientific notation (SCI) is different than the powers of 10 used in engineering (ENG)*. When converting to SCI powers of 10 from a ‘real’ number press the \[\text{SCI}\] button on your calculator, or put it in SCI mode and press the \[=\] key.

Wrap up: quote the number of H₂O molecules in 1 drop water using SCI notation:

1,000,000,000,000,000,000,000 molecules = ______________ molecules

REMEMBER: In chemistry we ALWAYS use scientific notation (SCI) for expressing large (>100) or small (<0.1) numbers.

________________________

Small Quantities

Question: How can very small numbers be expressed in SCI notation?

Just move the decimal point to the right until you get a single digit with decimals. The negative power of ten is the number of places the decimal point moved. Example:

\[
0.00125 = 1.25 \times 10^{-3} \quad \text{(number decimal places to right moved)}
\]

= ______________

Examples: Convert the following regular numbered quantities to powers of 10 (scientific notation). Try to do this without a calculator at first, then check with your calculator.
Review: You now know how to convert large or small ‘regular’ numbers into SCI notation either on paper or using your calculator.

Manipulating Large and Small Numbers

Enter the following SCI numbers into your calculator using the EE or EXP keys to express numbers in SCI notation, then press the = (in FLO mode) to obtain the ‘real’ number equivalent:

<table>
<thead>
<tr>
<th>‘Power of ten’ number (SCI)</th>
<th>Regular number (quantity) (FLO)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$5 \times 10^{-1}$</td>
<td></td>
</tr>
<tr>
<td>$1.5 \times 10^{3}$</td>
<td></td>
</tr>
<tr>
<td>$3.56 \times 10^{-3}$</td>
<td></td>
</tr>
</tbody>
</table>

Determine: $3 \times 10^7 \div 6 \times 10^3 = \underline{\phantom{000}}$

What answer did you get? Where any problems were encountered?
Making things even simpler – S.I. Prefixes

Certain powers of 10 can be replaced by a symbol known as a *decimal* (or *S.I.*) *prefix*

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>exa</td>
<td>E</td>
<td>1,000,000,000,000,000,000,000 (10^{18})</td>
</tr>
<tr>
<td>peta</td>
<td>P</td>
<td>1,000,000,000,000,000,000 (10^{15})</td>
</tr>
<tr>
<td>tera</td>
<td>T</td>
<td>1,000,000,000,000 (10^{12})</td>
</tr>
<tr>
<td>giga</td>
<td>G</td>
<td>1,000,000,000 (10^{9})</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>1,000,000 (10^{6})</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>1,000 (10^{3})</td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>0.1 (10^{-1})</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>0.01 (10^{-2})</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>0.001 (10^{-3})</td>
</tr>
<tr>
<td>micro</td>
<td>μ</td>
<td>0.000001 (10^{-6})</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>0.000000001 (10^{-9})</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>0.00000000000001 (10^{-12})</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>0.0000000000000001 (10^{-15})</td>
</tr>
<tr>
<td>atto</td>
<td>a</td>
<td>0.0000000000000000001 (10^{-18})</td>
</tr>
</tbody>
</table>
For decimal (S.I.) prefixes, just swap the appropriate “$x 10^n$” part of the number for the equivalent prefix’s symbol. Example:

$1.25 \times 10^{-3} \text{g} = 1.25 \text{mg}$ (milligrams)

**Task:** Convert the following quantities to SCI notation and decimal prefix notation:

<table>
<thead>
<tr>
<th>Quantity</th>
<th>With SCI notation</th>
<th>With Decimal Prefix</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0015 grams</td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.0000020 meters</td>
<td></td>
<td></td>
</tr>
<tr>
<td>30,000 dollars</td>
<td></td>
<td></td>
</tr>
<tr>
<td>12 million people</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Discussion:** Make a list of as many ‘everyday’ quantities as possible that use decimal prefixes (or similar related expressions):
**Significant Figures and Rounding Off**

**Question:** What are significant figures?

**Task:** Measure the length of your pencil (or some other object) in cm using a standard ruler. To how many sig. figs can you determine this value?

<table>
<thead>
<tr>
<th>Object Size measured (cm)</th>
<th>Number sig. figs.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Let’s figure out the rules for sig. figs. What is:

- 1.002 to 3 sig. figs.
- 1 sig. fig.
- 569.74 to 3 sig. figs.
- 4 sig. figs.
- 1 sig. fig.
- 0.00017 to 1 sig. fig.
Zeros before the first number are **NOT** counted as significant

Zeros after the first number **ARE** counted as significant

Round UP if the number after the last significant digit is > 5

Quote numbers in SCI notation if number sig. figs. < digits before decimal point.

Multiplication and Division (99% of your work is features one and/or the other*)

‘You are only as good as your worst measurement’

The result of any multiplication or division has the same number of sig. figs. as the measurement with the lowest number of sig. figs.

Example: Lead has a density of 11.4 g/cm³. What volume (in cm³) does 2.1 g of lead have?
Uncertainty in Measurement

‘Statistical’ Precision and Accuracy – group measurements

Discussion: What is the difference between precision and accuracy for a series of measurements? Can you be precise and at the same time inaccurate?

Think of a dartboard or other target when remembering the differences between precision (‘grouping’) and accuracy (‘hitting the bulls eye’)*

Notes

Good precision
Good accuracy
(Bull’s eye)

Notes

Good precision
Poor accuracy

Notes

Poor precision
Poor accuracy
This picture was found at a private individual’s internet site. The author states, among other things, that he “would make a bad murderer.”

How true, he fired 10 shots from 21 feet, (the target line was set at 75 feet) and hit the target 7 times.

How would you describe the spread of his shots in terms of accuracy and precision?

Formal definitions

**Accuracy**: How close individual measurements are to the ‘true’ (or ‘real’) value.

Accuracy is increased by repeating measurements (trials) and determining their average or mean. (Average value → true value for large number of trials)

**Precision**: How closely individual measurements agree with one another.

Standard deviation (or variance) is a relative measure of how precise a set of measurements are with respect to their average.

Precise measurements are not necessarily accurate – this is known as **systematic error**:

**Example**: Suppose a 2.0 mg grain of sand is accidentally dropped on to an analytical balance. Unless the balance is recalibrated, it will weigh ‘too heavy’ by 2.0 mg for each subsequent weighing. This is an example of a systematic (poor accuracy) error.

**Note**: Quantities determined from a series of measurements are most often quoted as their **mean ± standard deviation** for that data set. For large data sets (>20 measurements), this is equal to the **true value ± error**.
Individual measurements

For each individual measurement (mass, vol. etc.), its accuracy is typically quoted as to as good as the last measured (significant) digit. This fact is assumed for most physical measurements.

Example: For a mass reading of 21.245g, the last ‘5’ is uncertain so the accuracy of the reading is 21.245 ± 0.001 g.

Real Life: Dr. Mills owns a digital bathroom scale, which quotes weights to the nearest 0.1 pound (nice?). However, the scale often displays different weight values when a person steps off, then immediately steps back on to, the balance (not good), e.g. 212.4 lb, 212.7 lb, 212.5 lb.

How would you describe the performance of Dr. Mills’ scale, in terms of any individual measurement, with regard to accuracy and precision?
Temperature

**Background:** There are three temperature scales in common use today. Can you name them?

How were the end points of the two ‘metric’ scales defined? In other words, what natural conditions define these respective temperature values?

---

The Centigrade and Kelvin Scales

The Centigrade scale compared to the state of H₂O
Converting between Degrees Celsius and Kelvin

**Task:** By looking at the above figures, describe how the °C and K scales are related. What do they have in common? What is different?

1. Simply add 273.15 to ANY temp. quoted in °C to obtain the equivalent K value
   OR
   Simply subtract 273.15 from ANY temp. quoted in K to obtain the equivalent °C value

**Examples:**

1. What is 50°C in Kelvin?

2. What is 200 K in Celsius?
Comparing the Fahrenheit, Kelvin and Celsius Temperature Scales

Discussion: We saw that the end points for the °C scale corresponded to specific ‘natural’ temperatures – the same is true for the °F scale. What ‘natural’ temperatures do you think 0 °F and 100 °F (originally a centigrade scale) correspond to in nature. How about 212 °F and 32 °F?

Notes:

Diagram: Fahrenheit, Celsius and Kelvin thermometers side by side.
Task: By looking at the slide provided, describe how the °C and °F scales are related. What do they have in common? What is different?

These two basic differences between the °C and °F scales allow for equations relating them (conversion equations) to be constructed:

For converting °C to °F:

For converting °F to °C:

Task: Convert the following Celsius temperatures to °F and K

°F

K

Then ask me about the *extra credit* temperature….
Density

Review: How, in everyday words, is the property of density defined?

Density:

Where: ‘amount of matter’ = _______________

Discussion: What is the S.I. unit of density? Is this a convenient unit?

⇒ Density = _____________________________

Question: What are the two ‘convenient’ derived S.I. units of density used by chemists?

Ask me about the extra credit density….
Density Math

Recall: Density is defined by a simple equation, which has three related forms:

1. 
2. 
3. 

If you have problems with cross multiplication, remember that simple ‘pyramids’ can also be used to solve density and other 3 variable equations:

Example: 23.5 mL of a certain liquid weighs 35.062 g. What is the density of the liquid? What mass will 20mL of this liquid have?
Density of regular shaped objects

Regularly shaped objects (cubes, ‘bricks’, spheres, cylinders, cones…) have equations that define their volume.

Task: Sketch the following 3-D shapes and list the equations that define their volume (see your text book)

<table>
<thead>
<tr>
<th>Sketch of 3-D Shape</th>
<th>Volume equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cube</td>
<td>V =</td>
</tr>
<tr>
<td>‘Brick’</td>
<td>V =</td>
</tr>
<tr>
<td>Sphere</td>
<td>V =</td>
</tr>
<tr>
<td>Cylinder or disk</td>
<td>V =</td>
</tr>
</tbody>
</table>

1. Find the volume of the object in question via the equation that defines its volume (be sure to use cm for all length dimensions).

2. Substitute the derived volume value in \(D = \frac{M}{V}\) to find the object’s density (recall that mass is measured in grams).

Recall: the radius of a circle equals half of it’s diameter (i.e. \(\text{dia.} = 2r\))
Example: Dice used in Las Vegas weigh 2.65 g and have sides of length 1.2 cm. What is the density of a Las Vegas dice?

What relationship would you expect to exist between the density (macro), and associated physical state (macro), of any material with respect to spacing of its component particles (micro)? Exceptions??

### Densities of Some Common Materials

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Charcoal (from oak)</td>
<td>0.57</td>
</tr>
<tr>
<td>Ethanol</td>
<td>0.789</td>
</tr>
<tr>
<td>Ice</td>
<td>0.917 (at 0 °C)</td>
</tr>
<tr>
<td>Water</td>
<td>1.00 (at 4 °C)</td>
</tr>
<tr>
<td>Sugar (sucrose)</td>
<td>1.58</td>
</tr>
<tr>
<td>Table salt (sodium chloride)</td>
<td>2.16</td>
</tr>
<tr>
<td>Glass</td>
<td>2.6</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>2.70</td>
</tr>
<tr>
<td>Titanium</td>
<td>4.51</td>
</tr>
<tr>
<td>Iron</td>
<td>7.86</td>
</tr>
<tr>
<td>Copper</td>
<td>8.96</td>
</tr>
<tr>
<td>Lead</td>
<td>11.4</td>
</tr>
<tr>
<td>Mercury</td>
<td>13.55</td>
</tr>
<tr>
<td>Gold</td>
<td>19.3</td>
</tr>
<tr>
<td>Platinum</td>
<td>21.4</td>
</tr>
</tbody>
</table>
“Will it Float?”

The David Letterman Show on CBS often features a segment called ‘Will it Float’. Simply, Dave and Paul try to determine if an object, such as a refrigerator or 100 ft of insulation cable, will float when dropped into a large container of water.

Question: What physical property of a material will determine ‘if it will float’? What would be a more scientifically accurate (if less catchy) name for the ‘Will it float’ segment on Dave’s show?

Discussion: “Battleships and dating advice”

Task: Using the table supplied above, sketch a picture of what would happen if ~30 mL samples of ethanol, mercury and water, as well a lead ‘BB’ and a gold ring were added to a volumetric cylinder.
Dimensional Analysis (Conversion Factors)

Reading: Ch 1 section 8  
Homework: Chapter 1: 89*, 91, 93, 95, 99, 101, 119

* = ‘important’ homework question

Background

We do simple conversions between different units on a daily basis. For example:

**Question:**
How many eggs are there in 1 dozen eggs? _________

A statement such as this can be written as an identity

1 dozen eggs = 12 eggs

Recall: Is the above identity a measured or exact relationship? How would its use affect the number of significant figures used in any answer?

Using Identities and Conversion Factors

**Question:** How many eggs are there in 42 dozen eggs?

Use the appropriate identity to create a conversion factor. The conversion factor will transform the quantity into the desired form.

**Math:**

42 dozen eggs \( \times \) \( = \) ____________
Conversion factors are simply identities written as fractions. Each conversion factor has two ‘versions’

Task: Complete the following table by transforming the stated identity into its two corresponding conversion factors. Also include at least three additional conversion factors that you have encountered.

<table>
<thead>
<tr>
<th>Identity</th>
<th>Conversion factors</th>
<th>Exact? (Y/N)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 in = 2.54* cm</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 kg = 2.205 lb</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 m = 100 cm</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 ft = 12 inches</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Discussion:

How do you know which version of the conversion factor to use? Why?

For none exact identities and conversion factors, how many sig. figs are implied?

Example: Use the information from above to determine how many cm there are in 12.00 inches.

\[
12.00 \text{ inches} \times = _______________
\]

The unit belonging to the quantity and the denominator of the conversion factor cancel to leave a final answer with the desired unit

Generic Form:

Extra Credit (3 pts., typed for next time): Dr. Mills gets very cranky with the text as it says “1 in = 2.54 cm exactly”. What is the source of Dr. Mills’ confusion? Hint: Is the inch really an S.I. unit? Try searching for ‘International inch’ to get started.
Task: Complete the following conversions. See your text for appropriate conversion identities.

5.51 cm to meters

23.0 ounces to pounds

50.0 nm to meters*

6.56 miles to km

45.7 inches to cm

220 pounds to kg

# of drunk guys from 20 beers
Conversion Factor ‘Chains’

Question: How do you approach a problem like “Convert 55 cm into feet” – where there is no available single conversion factor?

Answer:

\[
55 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{3.281 \text{ ft}}{1 \text{ m}} = \text{ ____________}
\]

Link as many conversion factors as necessary together in order to create a ‘chain’. Each ‘link’ in the chain converts one unit to another and so on until the answer is reached.

Task: Complete the following ‘chain’ conversions. See your text for appropriate conversion identities.

4.00 ounces to grams

1.68 m to inches

5.8 km to feet
Question of the week: How many atoms would have to be placed end to end in order to reach Chicago. Assume an atom is 0.15 nm wide and Chicago is 40.0 miles away.

Conversion factors:

‘Chain’ Math:
The following questions were taken from your 1st practice midterm:

Question 1a (20 points) A spool of copper (Cu) wire has a mass of 2.00 pounds and a diameter of 50.0 μm. **Determine the wire’s mass, volume and length in the units specified below. Include any appropriate decimal prefixes in your final answers.**
Assume density copper (Cu) = 8.95 gcm$^{-3}$

Mass of the wire in kg:

(ANS = 0.907 kg)

Volume of the wire in cm$^3$:

(ANS = 101 cm$^3$)

Length of the wire in meters:

(ANS = 5.14 x 10$^4$ m)
Atomic Theory – Part 1

Reading: Ch 2 sections 1 – 6, 8  Homework: Chapter 2: 39, 47, 43, 49, 51*, 53, 55, 57, 71, 73, 77, 99, 103 (optional)

* = ‘important’ homework question

The Atomic Theory (John Dalton, 1803)

Dalton revisited the Ancient Greek Philosophers’ (Democritus et. al., 460 BC) ideas pertaining to how all matter is constructed from very small indivisible particles (‘atomos’).

Dalton formulated a set of ideas (postulates), known as “The Atomic Theory of Matter”, that would (~100 years) later be shown to be correct

Postulates of Dalton’s Atomic Theory of Matter

1. Matter is composed of extremely small particles called atoms

2. All atoms of a given element are identical, the atoms of each element are different and have different chemical and physical properties

3. Atoms are not changed into different types of atom(s) via chemical reactions. Atoms can neither be created nor destroyed

4. Compounds are formed when atoms of more than one type are combined. A compound always has the same relative number and kind of atoms

Notes on Dalton’s Atomic Theory
Atomic Structure

Discussion: In the introductory lectures we took a brief look at different types of matter (i.e. elements, compounds and mixtures). We know these materials are made from the smallest stable units of matter, atoms.

Atoms themselves are in turn made from smaller *unstable* particles – recall as many of these fundamental ‘building blocks’ of matter as you can:

**Question:** How are *all* of these fundamental ‘building blocks’ of matter related? Sketch a flow chart:

Fermi Lab, located in West Chicago, IL, is the world’s largest ‘atom smasher*’. Fermi is where scientists perform experiments in an attempt to understand the origins of the universe
Example: Water

In many ways we take atomic theory, as well as its eventual confirmatory experimental results, for granted (i.e. we don’t think about where it came from, we just assume and use it).

At the turn of the 19th century and for the next ~ 100 years this work was a the cutting edge of scientific research and was pursued by some of the world’s greatest scientific minds.

Task (complete outside of class): Complete the reading assignment for this note packet. Make notes on the following topics and make reference to the included illustrations in your discussions. This material, as well as other similar assignments, will likely form the basis of any extra credit questions appearing on midterm exams.
Cathode Rays and electrons (J.J. Thompson)

The Nuclear Atom (Rutherford)
Atoms and Isotopes

**Review:** Atoms are the smallest type of *stable* matter, they are typically spherical and have diameters of ~ 0.18 – 0.60 nanometers.

Shown on the left is an STM image of a silicon chip’s (Si (s)) surface. Note that it has a repeating ‘giant’ structure.

**Question:** Based on the scale, what is the approximate width of a silicon atom in nm?

**Answer:**

Ask me about the *extra credit* magnification….

The ‘Classical’ view of atomic structure

**Questions:**

1. What is found at the center of an atom?

2. What two different types of subatomic particle are found inside the nucleus? (*subatomic* means ‘smaller than’ atomic)

3. What ‘orbits’ the nucleus?
4. Sketch a generic diagram of an atom using the slide as a guide. Based on the slide, how many times smaller is the diameter of the nucleus than the atom as a whole?

Comparison of subatomic particles (i.e. the things atoms are made from)

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbol</th>
<th>Charge</th>
<th>Relative mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron</td>
<td>e</td>
<td>-1</td>
<td>1</td>
</tr>
<tr>
<td>Proton</td>
<td>p</td>
<td>+1</td>
<td>1836</td>
</tr>
<tr>
<td>Neutron*</td>
<td>n</td>
<td>0</td>
<td>1839</td>
</tr>
</tbody>
</table>

* ask me to tell you a very poor neutron joke - it starts with ‘a neutron walks into a bar’….

Electrons are much lighter than the neutrons and protons (that, in turn, ‘inhabit’ the nucleus) ⇒ ELECTRONS MOVE MUCH MORE QUICKLY THAN THE NEUCLIIUS EVER CAN (this is called the Born – Oppenheimer Approximation).

This is why electrons are said to either ‘orbit’ the nucleus or exist as ‘blurred out’ electron ‘clouds’. This is the main difference between the ‘classical’ and ‘modern’ models of atomic structure. We will study the modern ‘electron cloud’ model in depth later in the course.
Question: Are ATOMS* electrically charged?  
Answer: ______

Question: What then must be true for EVERY atom in terms of the number of electrons and protons it contains?

*Aside: We saw/will see in lab that ions are made by electrically charging atoms or molecules, we will study this concept in more detail later.

Question: Where can we find out the number of protons, Z, (and therefore also the number of electrons) an atom has?

**The Periodic Table**

Note how the P. Table is fundamentally arranged in terms of increasing atomic number (Z)
**Task:** Use the P. Table to determine how many protons and electrons the following types of atoms (elements) have:

<table>
<thead>
<tr>
<th>Atom</th>
<th>#p</th>
<th>#e</th>
<th>Atom</th>
<th>#p</th>
<th>#e</th>
<th>Atom</th>
<th>#p</th>
<th>#e</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon (C):</td>
<td></td>
<td></td>
<td>Silicon (Si):</td>
<td></td>
<td></td>
<td>Lead (Pb):</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Discussion:** The atomic number (Z) indicates the number of protons an individual atom has. What other type of subatomic particle is also found within an atom’s nucleus? How is the number of these particles within any nucleus represented and/or determined?
COMPLETE ATOMIC SYMBOL:

Example: Write the complete atomic symbol for an atom of Carbon that contains 6 protons and 6 neutrons.

‘Shorthand’ version of the complete atomic symbol:

Task: Carbon–14 has a mass number of 14. Use this information to write its complete atomic symbol. Do the same for U-235 and Cl-35.

* remind me to tell a story about U-235 and U-234
Understanding Isotopes

An element has a FIXED number of protons in its nucleus.
(This information is contained within the element’s Atomic Number. E.g. All hydrogen (H) atoms have 1 proton in their nuclei, while all carbon (C) atoms have 6 protons in their nuclei).

HOWEVER, an element can have a VARIABLE number of neutrons in its nuclei.
(This does NOT alter the identity of the element (#p same), but DOES make the element heavier or lighter (# n changed))

The AVERAGE atomic mass value for ALL an element’s isotopes is displayed in the periodic table.
E.g. Chlorine has a mass number of 35.45 amu* – there are NO single chlorine atoms in existence with a mass of 35.45 amu (i.e. no such thing as 0.45 of a neutron!), but there are Cl isotopes with mass numbers of 35 and 37 – their weighted average is 35.45 amu

The complete atomic symbol’s mass number’ (A) and the respective Element’s ‘box weight’ in the periodic table do NOT convey the same information.
The complete atomic symbol denotes the mass of ONE isotope of the element in amu, while the p. table gives is the average mass of ALL isotopes of the element in amu.

*Note: an amu is an atomic mass unit – the mass of a single proton or neutron. This is ≈1.66053873 x 10⁻²⁴ g.

It is much simpler to count atomic masses in amu – “an atom of carbon -12 (which contains 6 p and 6 n, so has a mass number of 12) weighs 12 amu” is better than saying “an atom of carbon -12 weigh 1.992648 x 10⁻²³ grams”!
Task: Complete the following table for the isotopes of Carbon. (Tip: what are the values of #p and #e ALWAYS for carbon? Where would you find this information?)

<table>
<thead>
<tr>
<th>Complete atomic Symbol</th>
<th>#p</th>
<th>#e</th>
<th>#n</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>5</td>
<td></td>
<td>7</td>
</tr>
<tr>
<td></td>
<td></td>
<td>6</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>8</td>
</tr>
</tbody>
</table>

I have a very poor $^{14}$C joke; ask at your own peril….

Determining Relative isotopic abundance

Typical Question: Naturally occurring magnesium has the following isotopic abundances:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Abundance</th>
<th>Mass (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{24}$Mg</td>
<td>78.99</td>
<td>23.98504</td>
</tr>
<tr>
<td>$^{25}$Mg</td>
<td>10.00</td>
<td>24.98584</td>
</tr>
<tr>
<td>$^{26}$Mg</td>
<td>11.01</td>
<td>25.98259</td>
</tr>
</tbody>
</table>

What is the average atomic mass of Mg?
The isotopic abundances of any series of isotopes *always* add up to 100%.

The sum of the *weighted abundances* is equal to the average atomic mass (as found in the P. Table).

Determine each isotope’s *weighted abundance* by multiplying its FRACTIONAL ABUNDANCE by its isotopic mass.

**Edit the table supplied to make two new columns - *fractional abundance* and *weighted abundance*.**

Determine the weighted abundances and then combine them to find the element’s average atomic mass.

**Answer:**

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Abundance</th>
<th>Fractional Abundance</th>
<th>Mass (amu)</th>
<th>Weighted Abundance (amu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{24}\text{Mg})</td>
<td>78.99</td>
<td>0.7899</td>
<td>x 23.98504</td>
<td>=</td>
</tr>
<tr>
<td>(^{25}\text{Mg})</td>
<td>10.00</td>
<td>0.1000</td>
<td>x 24.98584</td>
<td>=</td>
</tr>
<tr>
<td>(^{26}\text{Mg})</td>
<td>11.01</td>
<td>0.1101</td>
<td>x 25.98259</td>
<td>=</td>
</tr>
</tbody>
</table>

**Sum of weighted abundances** = _________________

**Check:** Is the sum of weighted abundances equal to the average atomic mass for Mg from the P. Table?
Task (complete outside of class): As the above table illustrates, the amu masses of individual atoms (isotopes), or even molecules, can be measured with a high degree of precision. This is made possible through the technique of mass spectroscopy. Make notes on the design and function of a mass spectrometer, as well as how the mass spectrum of Chlorine may be used to determine its average atomic mass (p 68, 69).
Write the **complete atomic symbol** for the isotope that contains 29 protons and 34 neutrons.

**Complete atomic symbol:**

**Answer:**

```
  63
 Cu
  29
```
Introduction to Molecules and Ions

<table>
<thead>
<tr>
<th>Reading: Ch 3 sections 1-6</th>
<th>Homework: Chapter 3: 23, 25, 27, 29*, 31, 33*, 35*, 39, 43, 49*, 51*, 53</th>
</tr>
</thead>
<tbody>
<tr>
<td>* = ‘important’ homework question</td>
<td></td>
</tr>
</tbody>
</table>

Molecules, Molecular Elements and Molecular Compounds

Recap: What is a molecule? What is a molecular compound? What is a molecular element?

| Molecule: |
| Molecular Element: |
| Molecular Compound: |

Molecules and their Chemical Formulas

There are two ways of describing the components (i.e. the number and type of atoms) found inside any molecule:

**Molecular Formula:** the actual number and type of atoms in a compound, e.g. hydrogen peroxide = H₂O₂

**Empirical Formula:** the lowest whole number ratio of each type of atom in a compound e.g. hydrogen peroxide = HO
Task: Complete the following table

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular formula</th>
<th>Empirical formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen peroxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dinitrogen tetroxide</td>
<td>N₂O₄</td>
<td></td>
</tr>
<tr>
<td>Benzene</td>
<td>C₆H₆</td>
<td></td>
</tr>
<tr>
<td>Butane</td>
<td>C₄H₁₀</td>
<td></td>
</tr>
<tr>
<td>Tetraphosphorus decoxide</td>
<td>P₄O₁₀</td>
<td></td>
</tr>
</tbody>
</table>

Note: Empirical formulas most often pertain to molecular / covalent compounds, as ionic compounds’ formulas are typically in their lowest ratio to begin with (this will be discussed further below)

Picturing Molecules – Structural Formulas

A structural formula is simply a more detailed version of the molecule’s corresponding molecular formula.

The major difference is that structural formulas also indicate the spatial relationship, and bonding, between atoms in a molecule

<table>
<thead>
<tr>
<th>Eg: Name and Molecular Formula</th>
<th>Electron density map of the ‘real’ molecule</th>
<th>Structural Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water (H₂O)</td>
<td><img src="image" alt="Water Molecule" /></td>
<td><img src="image" alt="Structural Formula" /></td>
</tr>
</tbody>
</table>
Task: Using the electron density maps as a guide, complete the following table

**Key**

<table>
<thead>
<tr>
<th>Key</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Hydrogen" /></td>
</tr>
<tr>
<td><img src="image" alt="Carbon" /></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Name and Molecular Formula</th>
<th>Electron density map of the ‘real’ molecule</th>
<th>Structural Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon Dioxide (CO₂)</td>
<td><img src="image" alt="Image" /></td>
<td></td>
</tr>
<tr>
<td>Methane (CH₄)</td>
<td><img src="image" alt="Image" /></td>
<td></td>
</tr>
<tr>
<td>Ethane (C₂H₆)</td>
<td><img src="image" alt="Image" /></td>
<td></td>
</tr>
<tr>
<td>Ammonia (NH₃)</td>
<td><img src="image" alt="Image" /></td>
<td></td>
</tr>
</tbody>
</table>
Naming Molecular Elements and Compounds

**Task:** Write the formula of *and* name as many molecular elements and compounds as you can

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Discussion:** What relationships do you see between the names and formulas of molecular compounds?
Prefix Table

<table>
<thead>
<tr>
<th>Number of atoms</th>
<th>Prefix*</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*Prefixes are dropped for the first single atom in a formula. E.g. CO₂ is named ‘Carbon dioxide’, not ‘Mono Carbon dioxide’.

Tasks:

Name the Following:  Write formulas for the following:

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NF₃</td>
<td>Chlorine dioxide</td>
</tr>
<tr>
<td>Cl₂O</td>
<td>Chlorine pentafluoride</td>
</tr>
<tr>
<td>P₂O₅</td>
<td>Dihydrogen monosulfide*</td>
</tr>
</tbody>
</table>

* If named using ionic nomenclature, also known as ________________
Location, Location, Location!

**ONLY** a non-metal bonded to another non-metal (top RHS p. table) make *molecular materials* with *covalent bonds*. E.g. CO, H_2O, SO_3

*These materials are named in accordance with the above ‘molecular’ rules*

---

**Metallic vs Non metallic Elements in the Periodic Table**

---

**ONLY** a non-metal (top RHS) bonded to metal (LHS) make *giant compounds with ionic bonds*. E.g. NaCl, CaO

*These materials are named in accordance with the ‘ionic’ rules discussed immediately below*
Ions and Ionic Compounds

Questions: What are ions? How are they made?

*Ask me to tell you a very poor ion joke.....

Atomic (micro) scale diagram of Ionization and macro scale crystal growth (slide)

In reality, electron(s) are EXCHANGED between atoms in order to become ionic compounds. I.E. what is lost by the metal (to become an $M^{n+}$ cation) is gained by the non-metal (to become $A^{n-}$ anion)
Making and Naming Ionic Formulas

List of Common atomic ions (must learn): See appendixes

<table>
<thead>
<tr>
<th>Group I</th>
<th>Group VII</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group II</td>
<td>Group VI</td>
</tr>
<tr>
<td>Group III</td>
<td>Group V</td>
</tr>
</tbody>
</table>

Naming atomic ions: An atomic (+ve) cation has the same name as the metal it was formed from. An atomic (-ve) anion has the same root as the non-metal it was formed from, but and –ide ending. Examples:

<table>
<thead>
<tr>
<th>Metal atom</th>
<th>Metal cation</th>
<th>Non-metal atom</th>
<th>Non-metal anion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td>O</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Ionic formulas are made by combining ANY cation (+ve) with any anion (-ve).

The order in ANY ionic formula is cation first, anion second, in both formula and name. i.e. (cation)(anion)

Examples: NaCl (sodium chloride)

LiF (sodium fluoride)
Ionic formulas ALWAYS have a ZERO net charge – i.e. the (+) and (-) ionic charges in ANY formula cancel.

If the above rule is followed, the ionic compound must exist and is probably sitting on a shelf in the chemistry stock room!

Task: Construct and name as many ionic compounds as possible from the following ions:

\[ \text{Li}^+ \quad \text{Ca}^{2+} \quad \text{Al}^{3+} \quad \text{Cl}^- \quad \text{O}^2- \quad \text{N}^{3-} \]

List of Common molecular ions (must learn): See attached handout.

Trick – many molecular ions appear on the data sheet (see handout). Just keep using (homework) and/or looking (fridge) at the rest

Naming molecular ions:

There is ONLY one molecular cation – (NH\(_4\))^+, ammonium.

Molecular anions with NO (or very few*) oxygen atoms in their structure have the –ide ending. Examples: 'OH (hydroxide)*, CN (cyanide)

Molecular anions with ‘lots’ of oxygen atoms in their structure have the –ate ending. Examples: (SO\(_4\))^2- (sulfate), (NO\(_3\))^− (nitrate), (CO\(_3\))^2- (carbonate), (PO\(_4\))^3- (phosphate)
**Recall:** Ionic formulas ALWAYS have a ZERO net charge – i.e. the ionic charges in ANY formula cancel.

This is true for molecular ions too – just treat the whole molecular ion as if it were an atomic ion when making the formula. Name the resulting compound in a similar way also.

**Task:** Construct and name as many ionic compounds as possible from the following ions:

- Li$^+$
- Mg$^{2+}$
- (NH$_4$)$^+$
- (NO$_3$)$^-$
- (SO$_4$)$^{2-}$
- (PO$_4$)$^{3-}$

**Naming Ionic compounds containing a cation of variable charge**

Metallic elements from the center of the periodic table (the *transition series*, between groups II and III) can form atomic ions with a range of +ve charges. Examples: Fe$^{2+}$ and Fe$^{3+}$, Cu$^+$ and Cu$^{2+}$.

**Question:** Can you see a potential problem with regard to writing the names and formulas of ionic compounds containing such cations?

**Answer:**
Ionic formulas featuring a variable charge (oxidation state) cation include the charge of the cation (written in Roman numerals) in the formula name. E.g.: Cu₂O = Copper(I) oxide

**Task:** Complete the following table:

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron (II) Sulfate</td>
<td></td>
<td>Copper (I) Phosphate</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Cu(NO₃)₂</td>
<td></td>
<td>FeCl₃</td>
</tr>
</tbody>
</table>

**Acids and bases**

**Discussion:** Are acids and bases typically ionic or molecular compounds (trick question!)? What is ‘special’ about them and their formulas?
Naming acids and bases: There are two ways of naming acids, and one way for bases:

1. Just use the standard approach for naming ionic compounds:

Remember:

- $\text{H}^+ = \text{‘hydrogen’ ion}$,
- $\text{OH}^- = \text{‘hydroxide’ ion}$.

Task: Name the following acids and bases using standard ionic compound nomenclature:

- $\text{HCl}$
- $\text{NaOH}$
- $\text{H}_2\text{SO}_4$
- $\text{Ca(OH)}_2$
- $\text{HNO}_3$
- $\text{Al(OH)}_3$

2. Using common nomenclature (chemical ‘nicknames’, must learn too)

Rules: Acids with $\text{–ide}$ anions (e.g. Chloride, $\text{Cl}^-$) have a ‘hydro’ prefix and an ‘–ic’ ending, followed by ‘acid’.

Example: $\text{HCl} = \text{Hydrochloric acid}$

Task: name the following acids:

- $\text{HBr}$
- $\text{HI}$
- $\text{HCN}$
- $\text{H}_2\text{S}$
Acids with molecular ‘ate’ anions, such as nitrate, (NO₃)⁻, and sulfate, (SO₄)²⁻, simply become ‘ic acids’:

Example: H(NO₃) = nitric acid

Task: name the following acids:

H₂SO₄  H₃PO₄

H₂CO₃
Ions in the Movies – Science fact or Science fiction?

Discussion: What makes for a good sci-fi movie? Why was Star Wars ‘good’ and Battlestar Galactic (released at the same time) ‘bad’??

‘Bad Guy’ Brian Cox

An ion cannon, as seen in The Empire Strikes Back has a lot in common with a computer technician’s static-guard wrist strap – electrical discharges can ‘fry’ sensitive electronics

Actual ion guns, used in surface science research and microchip manufacture.

Discussion: Would a commercially available ion gun be any use for ‘home defense’??
The following question were taken from your 1st practice midterm:

Write the formulas and names of nine ionic compounds that may be formed through combining the anions and cations ions listed immediately below.

H⁺ Cu²⁺ Al³⁺ Cl⁻ SO₄²⁻ PO₄³⁻

<table>
<thead>
<tr>
<th>Ionic Formula</th>
<th>Name of Ionic Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
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<td></td>
<td></td>
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<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Appendix: Table of Common Ions (Tro)

### Table 3.2 Metals Whose Charge Is Invariant from One Compound to Another

<table>
<thead>
<tr>
<th>Metal</th>
<th>Ion</th>
<th>Name</th>
<th>Group Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>Li⁺</td>
<td>Lithium</td>
<td>1A</td>
</tr>
<tr>
<td>Na</td>
<td>Na⁺</td>
<td>Sodium</td>
<td>1A</td>
</tr>
<tr>
<td>K</td>
<td>K⁺</td>
<td>Potassium</td>
<td>1A</td>
</tr>
<tr>
<td>Rb</td>
<td>Rb⁺</td>
<td>Rubidium</td>
<td>1A</td>
</tr>
<tr>
<td>Cs</td>
<td>Cs⁺</td>
<td>Cesium</td>
<td>1A</td>
</tr>
<tr>
<td>Be</td>
<td>Be²⁺</td>
<td>Beryllium</td>
<td>2A</td>
</tr>
<tr>
<td>Mg</td>
<td>Mg²⁺</td>
<td>Magnesium</td>
<td>2A</td>
</tr>
<tr>
<td>Ca</td>
<td>Ca²⁺</td>
<td>Calcium</td>
<td>2A</td>
</tr>
<tr>
<td>Sr</td>
<td>Sr²⁺</td>
<td>Strontium</td>
<td>2A</td>
</tr>
<tr>
<td>Ba</td>
<td>Ba²⁺</td>
<td>Barium</td>
<td>2A</td>
</tr>
<tr>
<td>Al</td>
<td>Al³⁺</td>
<td>Aluminum</td>
<td>3A</td>
</tr>
<tr>
<td>Zn</td>
<td>Zn²⁺</td>
<td>Zinc</td>
<td>*</td>
</tr>
<tr>
<td>Sc</td>
<td>Sc³⁺</td>
<td>Scandium</td>
<td>*</td>
</tr>
<tr>
<td>Ag²⁺</td>
<td>Ag⁺</td>
<td>Silver</td>
<td>*</td>
</tr>
</tbody>
</table>

### Table 3.3 Some Common Anions

<table>
<thead>
<tr>
<th>Nonmetal</th>
<th>Symbol for Ion</th>
<th>Base Name</th>
<th>Anion Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>F⁻</td>
<td>fluor</td>
<td>Fluoride</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl⁻</td>
<td>chlor</td>
<td>Chloride</td>
</tr>
<tr>
<td>Bromine</td>
<td>Br⁻</td>
<td>brom</td>
<td>Bromide</td>
</tr>
<tr>
<td>Iodine</td>
<td>I⁻</td>
<td>iod</td>
<td>Iodide</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O²⁻</td>
<td>ox</td>
<td>Oxide</td>
</tr>
<tr>
<td>Sulfur</td>
<td>S²⁻</td>
<td>sulf</td>
<td>Sulfide</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N³⁻</td>
<td>nitr</td>
<td>Nitride</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>P³⁻</td>
<td>phosph</td>
<td>Phosphide</td>
</tr>
</tbody>
</table>

### Table 3.4 Some Metals That Form Cations with Different Charges

<table>
<thead>
<tr>
<th>Metal</th>
<th>Ion</th>
<th>Name</th>
<th>Older Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chromium</td>
<td>Cr²⁺</td>
<td>Chromium(II)</td>
<td>Chromous</td>
</tr>
<tr>
<td></td>
<td>Cr³⁺</td>
<td>Chromium(III)</td>
<td>Chromic</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe²⁺</td>
<td>Iron(II)</td>
<td>Ferrous</td>
</tr>
<tr>
<td></td>
<td>Fe³⁺</td>
<td>Iron(III)</td>
<td>Ferric</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co²⁺</td>
<td>Cobalt(II)</td>
<td>Cobaltous</td>
</tr>
<tr>
<td></td>
<td>Co³⁺</td>
<td>Cobalt(III)</td>
<td>Cobaltic</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu⁺</td>
<td>Copper(I)</td>
<td>Cuprous</td>
</tr>
<tr>
<td></td>
<td>Cu²⁺</td>
<td>Copper(II)</td>
<td>Cupric</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn²⁺</td>
<td>Sn(II)</td>
<td>Stannous</td>
</tr>
<tr>
<td></td>
<td>Sn⁴⁺</td>
<td>Sn(IV)</td>
<td>Stannic</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg²⁺</td>
<td>Mercury(I)</td>
<td>Mercurous</td>
</tr>
<tr>
<td></td>
<td>Hg⁴⁺</td>
<td>Mercury(II)</td>
<td>Mercuric</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb²⁺</td>
<td>Lead(II)</td>
<td>Plumbous</td>
</tr>
<tr>
<td></td>
<td>Pb⁴⁺</td>
<td>Lead(IV)</td>
<td>Plumbic</td>
</tr>
</tbody>
</table>

### Table 3.5 Some Common Polyatomic Ions

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetate</td>
<td>C₂H₃O₂⁻</td>
<td>Hypochlorite</td>
<td>ClO₂⁻</td>
</tr>
<tr>
<td>Carbonate</td>
<td>CO₃²⁻</td>
<td>Chlorite</td>
<td>ClO₂⁻</td>
</tr>
<tr>
<td>Hydrogen carbonate (or bicarbonate)</td>
<td>HCO₃⁻</td>
<td>Chlorate</td>
<td>ClO₃⁻</td>
</tr>
<tr>
<td>Hydroxide</td>
<td>OH⁻</td>
<td>Perchlorate</td>
<td>ClO₄⁻</td>
</tr>
<tr>
<td>Nitrite</td>
<td>NO₂⁻</td>
<td>Permanganate</td>
<td>MnO₄⁻</td>
</tr>
<tr>
<td>Nitrate</td>
<td>NO₃⁻</td>
<td>Sulfite</td>
<td>SO₃²⁻</td>
</tr>
<tr>
<td>Chromate</td>
<td>CrO₄²⁻</td>
<td>Hydrogen sulfite (or bisulfite)</td>
<td>HSO₃⁻</td>
</tr>
<tr>
<td>Dichromate</td>
<td>Cr₂O₇²⁻</td>
<td>Sulfate</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>Phosphate</td>
<td>PO₄³⁻</td>
<td>Hydrogen sulfate (or bisulfate)</td>
<td>HSO₄⁻</td>
</tr>
<tr>
<td>Hydrogen phosphate</td>
<td>HPO₄²⁻</td>
<td>Cyanide</td>
<td>CN⁻</td>
</tr>
<tr>
<td>Dihydrogen phosphate</td>
<td>H₂PO₄⁻</td>
<td>Peroxide</td>
<td>O₂²⁻</td>
</tr>
<tr>
<td>Ammonium</td>
<td>NH₄⁺</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
# Chart of the Common Ions (Learn)

<table>
<thead>
<tr>
<th>+1 ions</th>
<th>+2 ions</th>
<th>+3 ions</th>
<th>-3 ions</th>
<th>-2 ions</th>
<th>-1 ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>H⁺</td>
<td>Mg²⁺</td>
<td>Al³⁺</td>
<td>N³⁻</td>
<td>O²⁻</td>
<td>F⁻</td>
</tr>
<tr>
<td>Li⁺</td>
<td>Ca²⁺</td>
<td>Fe³⁺</td>
<td>S²⁻</td>
<td>Cl⁻</td>
<td></td>
</tr>
<tr>
<td>Na⁺</td>
<td>Sr²⁺</td>
<td>Cr⁴⁺</td>
<td>PO₄³⁻</td>
<td>Br⁻</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>Ba²⁺</td>
<td>(phosphate)</td>
<td>SO₄²⁻</td>
<td>I⁻</td>
<td></td>
</tr>
<tr>
<td>Au⁺</td>
<td>Cu²⁺</td>
<td>(sulfate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ag⁺</td>
<td>Zn²⁺</td>
<td>CO₃²⁻</td>
<td>OH⁻ (hydroxide)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu⁺</td>
<td>Fe²⁺</td>
<td>(carbonate)</td>
<td>NO₃⁻ (nitrate)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₄⁺</td>
<td>Pb²⁺</td>
<td></td>
<td>CN⁻ (cyanide)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>(ammonium)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Solubility rules (will be supplied):

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Exceptions</th>
<th>Insoluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing NO₃⁻</td>
<td>None</td>
<td>Compounds containing CO₃²⁻</td>
<td>NH₄⁺ &amp; group IA cations</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>PO₄³⁻</td>
<td>NH₄⁺ &amp; group IA cations</td>
</tr>
<tr>
<td>Br⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>OH⁻</td>
<td>group IA cations</td>
</tr>
<tr>
<td>I⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>Ba²⁺, Sr²⁺, Ba²⁺ &amp; NH₄⁺</td>
<td></td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>Ba²⁺, Hg²⁺, Pb²⁺</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Tricks:**

Group # = ion charge for metal ions, e.g. Li (group 1) = +1

-(8 – group #) = ion charge for atomic non-metal ions, e.g. O = -(8-6) = -2

Formulas for most molecular ions appear on the solubility chart (supplied in data sheet).
Introduction to Stoichiometry

**Reading:** Ch 3 section 10  
**Homework:** Chapter 3: 93*, 95, 97*, 99*, 101

* = ‘important’ homework question

**Foundation:** The Atomic Theory (John Dalton, 1803):

3. Atoms are not changed into different types of atom(s) via chemical reactions. Atoms can neither be created nor destroyed

4. Compounds are formed when atoms of more than one type are combined. A compound always has the same relative number and kind of atoms

**Stoichiometry:** “The relationships among the quantities of reactants and products involved in a chemical reaction”.

Notes
Recall: A chemist’s goal is to explain *macroscopic* phenomena in terms of the repeated, identical reactions of its component *microscopic* particles. Recall the burning of charcoal example from week 1.

**Balanced chemical equations (micro scale) are used to describe larger (macro scale) events, e.g. The burning of charcoal (carbon) in air (contains oxygen gas) to produce carbon dioxide gas and heat:**

\[
C(s) + O_2(g) \rightarrow CO_2(g)
\]

This reaction is repeated many trillions of times when a charcoal brick is burnt in air to produce carbon dioxide gas and heat at a labor day ‘cook out’

---

What does each symbol represent in the above chemical reaction?

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Represents</th>
</tr>
</thead>
<tbody>
<tr>
<td>C(s)</td>
<td>A single atom of carbon in the charcoal brick (a reactant)</td>
</tr>
<tr>
<td>O_2(g)</td>
<td>A single molecule of oxygen gas from the air (a reactant)</td>
</tr>
<tr>
<td>→</td>
<td>A chemical symbol meaning ‘goes to’ or ‘is converted in to’</td>
</tr>
<tr>
<td>CO_2(g)</td>
<td>A single molecule of carbon dioxide gas produced from the reaction (the product)</td>
</tr>
</tbody>
</table>
Types of Chemical Reactions

Fact: There are FIVE general types of chemical reactions.

1. **Combination Reactions** - two or more types of material become one new material:

   **Generic:** \( A + Z \rightarrow AZ \)

   **Example:** \( C(s) + O_2(g) \rightarrow CO_2(g) \)

   **Note:** All combustion (adding oxygen) reactions are classed as combination reactions.

2. **Decomposition Reactions** - a material becomes two or more new materials:

   **Generic:** \( AZ \rightarrow A + Z \)

   **Example:** \( CaCO_3(s) \rightarrow CaO(s) + CO_2(g) \)

   **Note:** Decomposition reactions may be considered the reverse of combination reactions.

3. **Single Replacement (‘Prom’) reactions** - a more reactive material replaces a less reactive one in a compound:

   **Generic:** \( A + BZ \rightarrow AZ + B \)

   **Example:** \( Sn(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2(g) \)

   **Note:** The material replaced (B or H\(^+\) above) is said to be LESS reactive than its replacement (A or Sn above).
4. **Double Replacement reactions** - the respective ionic partners of a pair of dissolved ionic compounds are swapped, most often resulting in the formation of solid product(s):

\[
\text{Generic: } AX + BZ \rightarrow AZ + BX
\]

\[
\text{Example: } \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq)
\]

**Note:** These types of reactions typically take place between dissolved ionic compounds, and typically result in one of the new materials forming a solid precipitate (ppt)

5. **Neutralization reactions** - very similar to double replacement, but ALWAYS between an acid and a base:

\[
\text{Generic: } HA + MOH \rightarrow MA + HOH
\]

\[(\text{acid} + \text{base} \rightarrow \text{salt} + \text{water, H}_2\text{O})\]

\[
\text{Example: } \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)
\]

**Note:** These types of reactions are called neutralizations because acid (H\(^+\)) and basic (OH\(^-\)) ions react with each other to form water (H\(_2\)O). Such reactions typically liberate large amounts of heat (highly exothermic).
Task: Identify the following reactions as either: combination, decomposition, single replacement, double replacement or neutralization. Additionally, write the formula equivalent of each reaction below its word equation version.

sulfur(s) + oxygen gas → sulfur dioxide gas

magnesium carbonate(s) → magnesium oxide(s) + carbon dioxide gas

zinc(s) + copper (II) nitrate soln. → metallic copper + zinc (II) nitrate soln.

sodium hydroxide solution + hydrochloric acid solution →

Balancing Chemical Equations – Conservation of Mass Law

**Law of Conservation of Mass:** Atoms are neither created nor destroyed during a chemical reaction - they just change, add or lose partners:

\[ \text{e.g. } C(s) + O_2(g) \rightarrow CO_2(g) \]

There must (by definition) be the same number and type of each atom on both sides of the equation (i.e. before and after the reaction). The reactants and products must also have equal masses. Why?
Application: Rules for balancing chemical equations

1. Construct an unbalanced equation (sometimes using a word equation, as in lab). MAKE SURE THE FORMULAS ARE CORRECT, ONCE THEY ARE, DON’T ‘MESS’ WITH THEM!

‘Captain, I can ‘ney change the laws ‘o physics!’
....or chemistry either, Scotty – CORRECT CHEMICAL FORMULAS CAN NOT BE ALTERED ONCE WRITEN!

E.g.: Write an unbalanced equation using formulas for the combustion of magnesium (as seen in lab).

Magnesium metal (s) + Oxygen gas → Magnesium Oxide (s)

You know how to identify and write the name of either ionic or covalent (molecular) materials. What’s the trick?
⇒ Simply convert names to formulas first
2. Use balancing numbers (whole numbers that appear in front of formulas) to balance the equation: i.e. ensure that the law of conservation of mass is obeyed.

**UNDER NO CIRCUMSTANCES BE TEMPTED TO CHANGE THE FORMULAS!**

E.g.: Balance the following equation:

\[
\text{___ Mg (s) + ___ O}_2 \text{ (g)} \rightarrow ___ \text{MgO (s)}
\]

**Apply the ‘tennis’ approach:**
Balance a single type of atom on the left, then see what happens on the right; then balance a different type of atom on the right, then see what happens on the left etc. until done.

**Walk through:**
Balance # Mg atoms left and right
Balance # O atoms left and right
(re) Balance # Mg atoms left and right

**Micro scale view:**

\[
2 \text{ Mg} + \text{ O}_2 \rightarrow 2 \text{ MgO}
\]
Remember: As in this example, a chemist explains macro-scale phenomena in terms of a single (repeated many, many times...) microscopic description.
Task: Work in small groups to write balanced equations for the following.

1. The combustion of hydrogen gas to make water

2. The combustion of sodium metal to make its oxide

3. The dissolving of sodium metal in water to give sodium hydroxide and hydrogen gas.

4. The combustion of methane (CH$_4$).

The products of combustion for organic materials are carbon dioxide and water. Use the ‘tennis’ approach for balancing the C, H and then O atoms in order for organic (C, H and/or O containing) molecules.

$$\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)$$
5. The combustion of Ethanol (\( \text{CH}_3\text{CH}_2\text{OH} \))

6. The combustion of sugar (\( \text{C}_6\text{H}_{12}\text{O}_6 \))
Recap of skills learnt so far:

- Be able to write balanced chemical equation from a word description of the chemical process (this is what you are often required to do in lab)
- Be able to identify any reaction as belonging to one of the 5 general types of reaction - *combination*, *decomposition*, *single replacement*, *double replacement* or *neutralization*.

**Tricks**

1. **Construct an unbalanced equation** (sometimes from a word equation). MAKE SURE THE FORMULAS ARE CORRECT. ONCE THEY ARE, DON’T ‘MESS’ WITH THEM!

2. Use **balancing numbers** (whole numbers that appear IN FRONT of formulas) to balance the equation: i.e. ensure that the law of conservation of mass is obeyed.

   **Tip:** Apply the ‘tennis’ approach when balancing equations - balance a single type of atom on the left, then see what happens on the right; then balance a different type of atom on the right, then see what happens on the left etc. until done.

   **UNDER NO CIRCUMSTANCES BE TEMPTED TO CHANGE THE FORMULAS WHEN BALANCING AN EQUATION**

**Review Question:** Write a balanced reaction for the combustion of propane (C₃H₈(g)) in air. What general type of reaction is this?
“Balance”

Wrap up: Try the following questions (taken from the first practice midterm). See the next page for the answers

a. The burning of liquid butane (C₄H₁₀(l)) in air

b. The Neutralization of battery acid (sulfuric acid solution) with caustic soda (sodium hydroxide solution)

c. The reaction of solid diphosphorus pentoxide with water to form aqueous phosphoric acid

d. The decomposition of chalk (CaCO₃), when heated, to form solid calcium oxide and carbon dioxide gas

e. The reaction of metallic zinc with aqueous sulfuric acid to form aqueous zinc (II) sulfate and hydrogen gas
Answers:

a. The burning of liquid butane (\( \text{C}_4\text{H}_{10}(l) \)) in air

\[
2\text{C}_4\text{H}_{10}(l) + 13\text{O}_2(g) \rightarrow 8\text{CO}_2(g) + 10\text{H}_2\text{O}(g)
\]

b. The Neutralization of battery acid (sulfuric acid solution) with caustic soda (sodium hydroxide solution)

\[
\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})
\]

c. The reaction of solid diphosphorus pentoxide with water to form aqueous phosphoric acid

\[
\text{P}_2\text{O}_5(\text{s}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_3\text{PO}_4(\text{aq})
\]

d. The decomposition of chalk (\( \text{CaCO}_3 \)) when heated, to form solid calcium oxide and carbon dioxide gas

\[
\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})
\]

e. The reaction of metallic zinc with aqueous sulfuric acid to form aqueous zinc(II)sulfate and hydrogen gas

\[
\text{Zn(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})
\]
Quantitative Stoichiometry

<table>
<thead>
<tr>
<th>Reading:</th>
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<th>Homework:</th>
<th>Chapter 3: 55*, 57*, 59*, 63, 65*, 67, 75, 77, 79*, 81, 83</th>
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<tr>
<td>Ch 4 sections 1 – 4</td>
<td></td>
<td>Chapter 4: 25*, 27, 29, 31, 33*, 37*, 41*, 43, 47</td>
<td></td>
</tr>
</tbody>
</table>

* = ‘important’ homework question

Formula Weight

Question: What is the average mass (in amu) of 1 molecule of water (H₂O)? Where would you find the required information?

A molecule can be considered the ‘sum of its parts’ (atoms). Thus, simply add the masses of all the atoms in a single molecule to find the molecule’s mass (formula weight) in amu.

Determine formula weights for the following molecular materials:

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular Formula</th>
<th>Formula Weight (amu per molecule)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon dioxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂SO₄</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Diphosphorous pentoxide</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Percent Composition from Formulas

Analogy: If Dr. Phil weighs 250 pounds and his head alone weighs 25 pounds, then what % by mass (% composition or mass percent) is Dr. Phil’s head of his entire body?

Dr. Phil – a man with a relatively large head

\[
\% \text{ element} = \frac{(\text{number of atoms of element})(\text{atomic mass of element})}{\text{Formula weight of molecule}} \times 100\%
\]

Question: If one molecule of water has a mass of 18.02 amu, then what fraction of a water molecule’s mass is due to oxygen? What is this value as a percent?

Task: Determine the following:

1. The % mass of C and O in carbon dioxide

2. The % mass of C, H and O in formaldehyde (CH₂O).
The Mole

Just like the dozen, the MOLE is just a number that represents a bigger number. Since atoms and/or molecules are very small (i.e. to see a collection of atoms, say in your hand, you need a lot of them), the mole is a VERY large number:

1 dozen = 12 things (eggs)  1 mole = $6.02 \times 10^{23}$ things (e.g. atoms)

Task: To get an idea about how many atoms there are in a mole of atoms, write $6.02 \times 10^{23}$ as a regular number:

1 mole = __________________________________________________________________________

Note: The mole is sometimes called Avogadro’s number ($N_A$), so:

1 mole = $N_A = 6.02 \times 10^{23}$ things

Nerd stuff: When do you think some chemists celebrate mole day?

Discussion: Why is the mole ($N_A = 6.02 \times 10^{23}$) such a ‘strange’ number? Why not just $1 \times 10^{20}$ or something?
Proof: Determine how many carbon atoms (12.011 amu/atom) would weigh 12.011 grams. This number is, by definition, a mole ($N_A$) of C atoms (1 amu = 1.661 x 10^{-24} g).

Convert between grams, moles and number of atoms SEPARATELY using atomic mass values (Periodic table) and Avogadro’s number respectively.

\[
\text{Grams} \leftrightarrow \text{Moles} \leftrightarrow \text{Number of atoms}
\]

Task: Use the periodic table and the above information discussed above, determine the following quantities:

1. The mass of 2.0 moles of carbon atoms

2. The number of moles of lead (Pb) atoms in 35.5 grams of lead.

3. The number of Pb atoms in 35.5 grams of lead
Recall that a molecule is the ‘sum of its parts’ (atoms). Thus, simply add the masses of all the atoms in a molecule to find the molecule’s FW in amu, OR the mass of 1 mole of molecules (MOLAR MASS, \( \mathcal{M} \)) in grams. THIS IS THE POWER OF THE MOLE(!)

**Proof:** What is the mass in amu of 1 molecule of water (H\(_2\)O). What is the mass of one *mole* of water molecules in grams? (1 amu = 1.661 \( \times \) 10\(^{-24} \) g)

**Note:** The units of Molar mass (\( \mathcal{M} \)) are grams/mole (i.e. the number of grams in 1 mole of material).

**Task:** Calculate the Molar masses (\( \mathcal{M} \)) of the following compounds:

**Carbon dioxide:**

**Diphosphorous pentoxide:**

**Calcium chloride:**
Just like with atomic masses, Molar masses can be used to convert between grams, moles and number of molecules.

Grams ↔ Moles ↔ Number of molecules

Worked Example: How many molecules of sugar (C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}) are there in a 2.15 gram packet of sugar?

Plan: Write down what you are given and what you can immediately figure out:

Mass sugar = 2.15 g

\[
\text{Mass} \text{ C}_6\text{H}_{12}\text{O}_6 = 6\text{C} + 12\text{ H} + 6\text{ O}
\]

\[
= 6(12.011 \text{ g/mol}) + 12(1.01\text{ g/mol}) + 6(16.00 \text{ g/mol})
\]

\[
= 180.2 \text{ g/mol}
\]

Use the conversion factor of \[1 \text{ mol C}_6\text{H}_{12}\text{O}_6 = 180.2 \text{ g}\] to find # moles sugar in 2.15 g of sugar:

\[
2.15 \text{ g} \times \frac{1\text{ mol}}{180.2 \text{ g}} = 0.0119 \text{ mols}
\]

Use the conversion factor of \[1 \text{ mol} = 6.02 \times 10^{23} \text{ molecules}\] to find # molecules of sugar in 0.0119 mols. (2.15 g) of sugar:

\[
0.0119 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 7.16 \times 10^{21} \text{ molecules}
\]

Remember: You CANNOT convert directly from grams to number of molecules. i.e. YOU MUST ALWAYS GO THROUGH MOLES:

Grams ↔ Moles ↔ Number of molecules
Things to remember:

Key relationships:

FW = sum of all atomic masses in a SINGLE molecular formula (amu/molecule)

\( \mu = \text{sum of all atomic masses in any SINGLE molecular or ionic formula (grams/mole)} \)

\[ \mu = \frac{\text{number grams material}}{\text{number moles of material}} \]

1 mole = 6.02 x10^{23} particles

FW and \( \mu \) have identical numerical values but DIFFERENT units

You can write a conversion pyramid showing the relationship between \( \mu \), g and mols!

The relationship between \( \mu \), #grams and #moles is THE most frequently used equations in chemistry? Why?

Answer:

Observation(!):
Task: Determine the following quantities:

1. The number of molecules in 5.0 g of oxygen gas

2. The number of molecules in 12.5 grams of P_2O_5 (s)

3. The mass in grams of 2.76 \times 10^{23} molecules of H_2O (l)
Empirical & Molecular Formulas from Analyses

Recall: Mass Spectrometry

Scientists are able to determine the mass of *individual* atoms and/or molecules using a *mass spectrometer*. The resulting mass spectrum gives a parent peak (M⁺), which gives (is) the *formula weight* of the specie under study.

Ever wondered why the masses of individual isotopes for any stable element *or molecule* are known to a large number of sig. figs? They are measured using a very accurate / precise mass spectrometer.

**Extra Credit (to be completed outside of class):** Review the mass spec. section on page 68 of your text *once more*. Use this information to determine the identities of the molecular fragments giving rise to each peak in the ‘mass spec’ of water.

Hand in a concise, typed report next lecture for 2 pts e.c.
Determining Empirical and Molecular Formulas ‘by hand’

Recall the definitions of molecular formula and empirical formula:

**Molecular Formula**: the actual number and type of atoms in a compound, e.g. hydrogen peroxide = H₂O₂

**Empirical Formula**: the lowest whole number ratio of each type of atom in a compound e.g. hydrogen peroxide = HO

Recall that the Empirical formula is the lowest whole number ratio of each type of atom in a compound ⇒ find the moles of each type of atom and then find their ratio.

**Worked Example**: A 1.271 g sample of Al(s) was allowed to react with chlorine. The mass of aluminum chloride produced was 6.280 g. Determine the empirical formula of aluminum chloride.

**Step 1.** Write an unbalanced chemical equation (do not assume balancing numbers or formulas for these problems). Find the mass of the missing reactant by applying the conservation of mass law.

\[ \text{Al} + \text{Cl} \rightarrow \text{Al}_x\text{Cl}_y \]

1.271 g 6.280 g

**Step 2.** Find the moles of each reactant using the atomic masses from the periodic table.
Moles Al =

Moles Cl =

**Step 3.** Substitute the # moles determined for each type of atom in the product’s empirical formula.

\[
\text{Al} \quad \text{Cl}
\]

**Step 4.** Find the lowest whole number ratio of each type of atom in the empirical formula. This is the final answer.

\[
\text{Al} \quad \text{Cl}
\]
Finding the Molecular Formula from the Empirical Formula

Recall that the molecular formula is some whole number of times larger than the empirical formula (e.g. H₂O₂ compared to HO (x2)).

⇒ the molecular formula will be ‘heavier’ than the empirical formula by the same factor (x2)

Mass spectroscopy data (supplied) will give the Formula Weight (FW) of the element or compound under study. This value is then compared with the empirical formula weight (EW) to find the molecular formula

Task: Work out the molecular masses of H₂O₂ and HO. What is their ratio?

\[ M_{H_2O_2} = \quad M_{HO} = \]

Ratio =

Find the ratio of the molecular formula to the empirical formula – this information tells you how much ‘bigger’ the molecular formula is than the empirical formula.
**Worked Example:** Ethylene glycol contains 38.7% C, 9.7% H and 51.6% O. Calculate the empirical and molecular formula of ethylene glycol given its molar mass = 60 g/mol (60 amu/molecule)

When given the % by mass values of each atom in a compound assume a 100 g sample – the % and g values are then the same.
‘Slides and Ladders’ – Finding Theoretical Yield and % Yield

Important Definitions

Theoretical Yield: The amount of product, in grams, expected (calculated) for a reaction.

Actual Yield: The amount of product, in grams, recovered (weighed) for a reaction.

\[
% \text{ Yield} : \quad \% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%
\]

Discussion: Are theoretical and actual yields ever the same (i.e. does % yield = 100%) in practice? What factors influence the % yield?

Finding the Theoretical Yield using ‘Slides and Ladders’

Worked Example: If 5.00 g of Propane (C\textsubscript{3}H\textsubscript{8}(l)) is combusted in excess oxygen gas, what mass of water is expected to be formed? What mass and volume of CO\textsubscript{2} (g) (at STP) would you expect to collect?

Step 1. Write a balanced Chemical Equation and
Step 2. Set up g, \( M \) and mole ‘ladder’ grid

Step 3. Fill in the ladder grid with as much information as possible – this is typically supplied gram weights and molar mass data.

Step 4. Convert g \( \rightarrow \) moles by ‘climbing’ down ladder(s) (g / \( M \) = moles).

\[ \text{Remember: Molar masses are calculated for ONE molecular formula only. I.E. ignore any balancing numbers when figuring out } M \text{ values} \]

Step 5. Convert moles reactant \( \rightarrow \) moles product(s) by comparing balancing numbers and ‘sliding’ across.

Step 6. Convert moles product(s) \( \rightarrow \) grams product(s) by ‘climbing’ up ladder(s) (moles x \( M \) = grams).

Note: moles of gas can be converted to Liters of gas using:

\[ 1 \text{ mole any gas} = 22.4 \text{ L at STP} \]
Group Task: Determine the following quantities

1. What mass of dissolved HCl is needed to completely react 5.00g of CaCO$_3$(s), according to the following *unbalanced* reaction? What volume of CO$_2$(g) is generated at STP?

\[ \text{CaCO}_3(\text{s}) + \_\_ \text{HCl(aq)} \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O(l)} \]

2. What mass of magnesium oxide is recovered when 1.56 g of Mg(s) is burnt in air to give MgO(s)? What volume of oxygen gas is consumed during this process (assume STP).

3. A student recovers 1.59 g of CaCO$_3$(s) from an experiment when they should have produced 2.32 g. What is the student’s % yield for their reaction?

4. Let’s talk about your ‘Alum’ lab.....
Recall: Reactants ALWAYS combine in the ratio defined by their respective balancing numbers:

\[ 1 \text{ AgNO}_3(\text{aq}) + 1 \text{ NaCl}(\text{aq}) \rightarrow 1 \text{ AgCl}(s) + 1 \text{ NaNO}_3(\text{aq}) \]

i.e. 1 mole of AgNO\(_3\)(aq) will react exactly with 1 mole of NaCl(aq)

Problem: It is VERY difficult to add exactly the right ratio of reactants in the lab ⇒

There will be too much of one reactant (the excess (XS) reactant)
There will be too little of one reactant (the limiting reactant)

Discussion questions: If 1.0 moles of AgNO\(_3\)(aq) is added to 1.5 moles of NaCl(aq), then:

1. Which reactant is INXS?
2. Which reactant is limiting?
3. How many moles of AgCl(s) would be formed?
4. How many moles of NaCl(aq) would remain unreacted?

Compare the ‘Ideal’ (from the balanced equation) and the ‘Real’ (given) ratio of reactants to determine which is the limiting reactant (AgNO\(_3\)(aq))

Since the LIMITING reactant will run out first, it determines the amount of product that can be formed (as well as the amount of XS reactant that is left behind)
**Questions:** Work out the mass of each product formed in the following reactions, assuming 10.0 grams of each reactant are initially mixed together.

Use a regular slides and ladders approach, but ‘slide’ across (to find moles of product) using the molar ratio determined by the limiting reactant.
Use conversions factors to find out the number of moles required to react with one of the reactants in ‘NON 1:1 problems’

1. \( \text{CaCO}_3(s) + \underline{\text{HCl}}(aq) \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \)

2. \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq) \)

3. \( \underline{\text{Al}}(s) + \underline{\text{Cl}}_2(\text{g}) \rightarrow \underline{\text{AlCl}}_3(s) \)
Try the following questions (taken from the first and second practice midterms).

“% Mass”

Calculate the % mass of each type of atom in the following materials:

$\text{BaSO}_4$

Ba = 58.8%
S = 13.8%
O = 27.4%

$\text{NO}$

N = 46.7%
O = 53.3%
"Limiting"

Lithium and nitrogen react to produce lithium nitride:

\[ 6 \text{Li(s)} + \text{N}_2 (g) \rightarrow 2 \text{Li}_3\text{N} (s) \]

If 7.00 grams of each reactant undergo a reaction with 85.0% yield, how many grams of Li$_3$N are obtained from the reaction?

[10.0g]
“Take out?”

Monosodium glutamate (MSG) is a popular flavor enhancer used in many fast foods. Assuming MSG contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.6% Na, and has a molar mass of 169 g/mol determine:

The Empirical formula of MSG

\[[\text{C}_5\text{H}_8\text{O}_4\text{NNa}]\]

The Molecular formula of MSG

\[[\text{C}_5\text{H}_8\text{O}_4\text{NNa}]\]
Aqueous Solutions & Reactions

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<th>Homework:</th>
<th>Chapter 4: 49*, 53, 55, 57, 63, 67, 69*, 71*, 73, 75*</th>
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</thead>
</table>

* = ‘important’ homework question

Concentration of Solutions

Discussion: When you finish this class chances are good you’ll head out to a local bar for a well deserved ‘adult beverage’. Assuming the bar you visit is running a special where you can buy a pint of beer or a pint of wine for the same price, which do you choose and why?

Answer:

Concentration = Molarity* (M) = number moles of solute per Liter

of solution

* The Molarity (mol/L conc.) is most often printed on lab reagent solution bottles, e.g. 9.0 M H₂SO₄ (aq)

i.e. Molarity = \[
\text{Moles Solute} \over \text{Liters Solution}
\]

Units: mol/L or just M

Where:

\[\text{SOLUTION} = \text{SOLUTE} + \text{SOLVENT}\]

Example:

PEPSI =

Task: Think up one more example illustrating the components of a solution.
Stock Solutions: It is often necessary to make up a stock solution (a solution of known concentration) in lab for use within an experimental procedure – you will do this with your titration labs.

Walkthrough: Preparing a 1.0 M NaCl (aq) stock solution:

1. Weigh out and add 1.00 mol of NaCl.
2. Add water until solid is dissolved. Then add additional water until the 1-liter mark is reached.
3. Mix.

A 1.00 molar NaCl solution

Notes
Question and Demo: What is the concentration of a solution made by dissolving 2.845 g of sugar (C₆H₁₂O₆) in water and making the final volume up to 150 mL?

Remember that Moles Solute, Solution Concentration and Solution Volume are all related. Sketch a ‘pyramid’ to help you solve concentration problems for solutions.

Question: What mass of NaCl is contained within 0.50 L of a 6.0 M NaCl(aq) solution?
Since \( CV = \text{moles} \) for any solution, concentration (C) and solution volume (V) terms can also be used in ‘slides and ladders’ problems featuring solutions.

Example: What volume (in mL) of 2.0 M HCl is needed to completely react with a 5.5 gram stick of chalk (CaCO\(_3\)(s))? What volume of CO\(_2\)(g) would be recovered, assuming a 100% yield? **Hint:** You have seen this reaction before.
Dilutions

**Question:** When you dilute something with water (for example your morning coffee) do you change the number of moles of solute (caffeine) dissolved?

**Answer:**

Since the number of moles of solute in a solution is not altered by adding more solvent (by dilution) then:

\[
\text{Moles solute (before dilution)} = \text{Moles solute (after dilution)}
\]

Now, since moles is a constant and moles = CV, then:

\[
C_{(\text{before})}V_{(\text{before})} = C_{(\text{after})}V_{(\text{after})}
\]

Use the \(C_{(\text{before})}V_{(\text{before})} = C_{(\text{after})}V_{(\text{after})}\) relationship to find the concentration of a solution that has been diluted

**Example Question:** A laboratory procedure requires you to prepare 100 mL of 2.0 M HCl solution from an 18.0 M stock solution of this acid. Describe how this would be achieved.
Recap: The Importance of the mole

Most of the equations we have met in this handout feature moles as a variable. Thus, moles can in many ways be considered the chemists’ link between macro and micro scale quantities.

Task: Write down as many equations you can featuring the mole. Use this information to construct a ‘spider’ (flow) chart illustrating how all these conversions ‘go through’ moles.
Properties of Solutions: Electrolytes

An electrolytic solution is formed when an ionic compound (such as NaCl) dissolves in water to form a solution containing solvated (dissociated and mobile) ions. Alternatively, an electrolyte is simply a dissociated ion dissolved in water. (see below and slide)

Example: When table salt is dissolved in water it undergoes dissociation to make aqueous Na$^+$ and Cl$^-$ ions (this is what happens when you dissolve table salt to make pasta!):

$$\text{NaCl (s)} \rightarrow \text{Na}^+ (aq) + \text{Cl}^- (aq)$$

Note: This information is most often further abbreviated, i.e.: $\text{NaCl (aq)} \equiv \text{Na}^+ (aq) + \text{Cl}^- (aq)$

This is true for all electrolytic solutions. Recall that the state symbol (aq) means ‘aqueous’ or with water – it means the solute has formed a solution by dissolving in a solvent (water).
Aqueous Reactions in which a solid product (precipitate, ppt.) forms

Worked example: Write balanced, complete and net ionic equations for the reaction of silver nitrate and sodium chloride solutions.

The Balanced ionic equation

This is the balanced, ionic double replacement (‘swapping partners’) reaction you are already familiar with.

Balanced:

Question: How do we know which one of the products formed (AgCl or NaNO₃) is the solid recovered from the reaction?

Whether or not an ionic material is soluble in water is known. This data appears in the solubility chart.

⇒ Determine if the product(s) are soluble via the chart and include the appropriate state symbol(s) in the balanced equation. ALWAYS INCLUDE STATE SYMBOLS IN ALL YOUR BALANCED, COMPLETE AND NET EQUATIONS

Solubility chart (will be supplied)

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Exceptions</th>
<th>Insoluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing</td>
<td>NO₃⁻</td>
<td>None</td>
<td>CO₃²⁻</td>
</tr>
<tr>
<td></td>
<td>Cl⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>PO₄³⁻</td>
</tr>
<tr>
<td></td>
<td>Br⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>OH⁻</td>
</tr>
<tr>
<td></td>
<td>I⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>Ca²⁺, Sr²⁺, Ba²⁺ &amp; NH₄⁺</td>
</tr>
</tbody>
</table>
Task: State whether the following ionic materials are soluble or insoluble in water:

<table>
<thead>
<tr>
<th>Material</th>
<th>Material</th>
<th>Material</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaNO₃</td>
<td>PbCl₂</td>
<td>PbSO₄</td>
</tr>
<tr>
<td>AgCl</td>
<td>Ca(OH)₂</td>
<td>AgOH</td>
</tr>
</tbody>
</table>

The Complete ionic equation

Recall that the state symbol (aq) means that the ions of an ionic compound are SEPARATED (dissociated) and, therefore, MOBILE when dissolved in water.

A complete ionic equation shows each type of ion dissolved in solution. Consider it to be a ‘longhand’ version of the balanced equation.

Balanced:

\[
\text{AgNO}_3 (aq) + \text{NaCl}(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3 (aq)
\]

Complete:

Solids (s), liquids (l) and gases (g) are NOT in the (aq), state so do not appear as individual (dissociated) ions in the complete ionic equation.
The Net ionic equation

**Question:** What can be done ‘algebraically’ with the complete ionic equation?

**Complete:**

\[ \text{Ag}^+_{(aq)} + \text{NO}_3^-_{(aq)} + \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} \rightarrow \text{AgCl}_{(s)} + \text{Na}^+_{(aq)} + \text{NO}_3^-_{(aq)} \]

**Net:**

‘Cancel’ ions (spectator ions) common to both sides in the Complete ionic equation to give the Net ionic equation. This is ‘where the action is’ in any ionic reaction.

**Walkthrough Example:** Balanced, complete and net ionic equations for the reaction of potassium iodide and lead (II) nitrate solutions:

**Balanced**

\[ \_\_\_ \text{Pb(NO}_3\text{)}_{2}\_\_\_ \text{ (aq)} + \_\_\_ \text{KI (aq)} \rightarrow \_\_\_ + \_\_\_ \]

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Complete

KI (aq) and Pb(NO₃)₂ (aq) → Pbi₂ (s) and KNO₃ (aq)

Net

Summary

2 KI(aq) (soluble) + Pb(NO₃)₂(aq) (soluble) → 2 KNO₃(aq) (soluble) + Pbi₂(s) (insoluble)
Group Task: Write balanced, complete and net ionic equations for the following reactions:

1. Silver nitrate solution mixed with calcium hydroxide solution

2. Barium nitrate solution mixed with sulfuric acid solution

3. Calcium chloride solution mixed with sodium phosphate solution (lab)
"Equations and Solubility"

Try the following question (taken from the second practice midterm).

Write balanced, complete and net ionic equations illustrating the reaction between aqueous solutions of silver nitrate and sodium phosphate. Include all state symbols.

Balanced chemical equation:

Complete ionic equation:

Net ionic equation:

List the names and formulas of five insoluble ionic compounds containing the hydroxide ion.
"Equations and Solubility"

Try the following question (taken from the second practice midterm).

Write *balanced*, *complete* and *net ionic* equations illustrating the reaction between aqueous solutions of silver nitrate and sodium phosphate. *Include all state symbols.*

**Balanced chemical equation:**

\[
3 \text{AgNO}_3 (aq) + \text{Na}_3\text{PO}_4 (aq) \rightarrow \text{Ag}_3\text{PO}_4 (s) + 3 \text{NaNO}_3 (aq)
\]

**Complete ionic equation:**

\[
3\text{Ag}^+ (aq) + 3\text{NO}_3^- (aq) + 3\text{Na}^+ (aq) + \text{PO}_4^{3-} (aq) \rightarrow \text{Ag}_3\text{PO}_4 (s) + 3\text{Na}^+ (aq) +3\text{NO}_3^- (aq)
\]

**Net ionic equation:**

\[
3\text{Ag}^+ (aq) + \text{PO}_4^{3-} (aq) \rightarrow \text{Ag}_3\text{PO}_4 (s)
\]

List the names and formulas of five insoluble ionic compounds containing the hydroxide ion.

Any except with group 1A, Ca$^{2+}$, Sr$^{2+}$, Ba$^{2+}$ and NH$_4^+$

eg: Silver (I) hydroxide AgOH, Copper (I) Hydroxide CuOH, Magnesium hydroxide Mg(OH)$_2$, Iron (II) hydroxide Fe(OH)$_2$, Lead (II) hydroxide Pb(OH)$_2$


Acid – Base Reactions

Reading: Ch 4 section 8  
Homework: Chapter 4: 79, 81*, 83*, 108 (optional)

* = ‘important’ homework question

Background

Recall: Acids and bases are ‘special’ in two ways:

What is special about their formulas?

Acids:  
Bases:

What is special about the state of matter they (typically acids) exist as in the:
(a) pure and (b) aqueous (dissolved) states?

Pure:  
Dissolved (aq):

Task: Write formulas for the following common lab reagents:

Hydrochloric acid  
Sodium hydroxide

Nitric acid  
Calcium hydroxide

Sulfuric acid  
Potassium hydroxide

Assuming each of the above is of the same concentration, which acid and which base are the ‘strongest’? Why?
Monoprotic acids dissociate to give 1 $\text{H}^+$ (aq) ion per formula unit
E.g.: HCl, HNO$_3$

Diprotic acids dissociate to give 2 $\text{H}^+$ (aq) ions per formula unit
E.g. H$_2$SO$_4$

Details: Polyprotic acids undergo *sequential dissociation* when dissolved in water. For Sulfuric acid:

\[
\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{H}^+ (\text{aq}) + \text{HSO}_4^- (\text{aq})
\]

\[
\text{HSO}_4^- (\text{aq}) \rightarrow \text{H}^+ (\text{aq}) + \text{SO}_4^{2-} (\text{aq})
\]

(Combine the equations to obtain the overall reaction for the dissolution of Sulfuric acid)

Similarly....

Monobasic alkalis dissociate to give 1 OH$^-$ (aq) ion per formula unit
E.g.: NaOH, KOH

Dibasic alkalis dissociate to give 2 OH$^-$ (aq) ions per formula unit
E.g. Ca(OH)$_2$

Example: The dissociation of calcium hydroxide
**Strong and Weak Acids and Bases** (recall previous packet)

*Strong (mineral) acids* undergo complete (~100%) dissociation, so give a high conc. of \(H^+\)(aq) ions in solution. E.g. HCl, HNO₃

*Weak (organic) acids* undergo partial (<<100%) dissociation, so give a low conc. of \(H^+\)(aq) ions in solution E.g. H(C₂H₃O₂)

**Discussion**: A student measured the pH of equal volumes of 0.5 M HCl (hydrochloric acid) and 0.5M HC₂H₃O₂ (acetic acid). Which solution would you expect to have the lower pH (i.e. the highest conc. of \(H^+\) (aq) in solution)? Hint: Recall analogous ICF tables for strong and weak electrolytes.

![HCl (aq) and HC₂H₃O₂ (aq) images]

Strong (mineral) acids and bases give the best ‘bang for the buck’ (most \(H^+\)(aq) or \(OH^-\)(aq) ions respectively in solution per formula unit), so are used as lab reagents.
OK, what about weak bases then….

Case study: The weak base, aqueous ammonia (NH₃ (aq))

Background: You may have noticed in lab that aqueous ammonia is used as a ‘gentle’ base, (i.e. it is a weak base). How can this be? Can NH₃ (aq) dissociate to give OH⁻ (aq) ions??

Answer:

Explanation:

Equation:

Task: Write a ICF table for a solution of NH₃ (aq) that is ‘10% basic’

<table>
<thead>
<tr>
<th>Eqn.</th>
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<tbody>
<tr>
<td>I</td>
<td></td>
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<tr>
<td>C</td>
<td></td>
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<tr>
<td>F</td>
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</tbody>
</table>
Neutralization Reactions

Recall: Acids and bases react with each other (often via a titration) in a ‘special’ double replacement reaction called a neutralization reaction.

Word Equation:

Generic Equation:

Task: Write balanced, complete and net ionic equations for the reaction of hydrochloric acid solution with sodium hydroxide solution.
Summary

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.
Review of Titrations

In practice, an *acid soln* (eg. HCl) is placed in a flask, with a few drops of an *indicator*, and then *titrated* with a *base soln* (eg. NaOH) via a buret:

The equivalence point of any acid plus base titration (reaction) is reached when:

\[
\text{Moles acid (H}^+\text{) = Moles base (OH}^-\text{)}
\]

i.e. when (for a pair of *mono* reagents):

\[
C_{\text{acid}} \times V_{\text{acid}} = C_{\text{base}} \times V_{\text{base}}
\]
The indicator in an acid base titration changes color at the equivalence point – i.e. where the moles of acid (H⁺) is equivalent to the moles base (OH⁻) added – this is when you stop the buret!!

Titrations are just a specific application of slides and ladders – the reactants are always an acid and a base (while the products are always a salt and water)

Example: What volume of 0.128 M HCl (aq) is needed to neutralize 2.87 g of Mg(OH)₂? Careful! – the alkali is dibasic.

\[
\_ \text{HCl (aq)} \ + \ _\_ \text{Mg(OH)}_2 (s) \ + \ _\_ \text{MgCl}_2 (aq) \ + \ _\_ \text{H}_2\text{O (l)}
\]

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**Acid plus Carbonate Reactions**

An acid plus carbonate reactions always produce a salt, water and carbon dioxide gas. The word equation is:

\[
\text{ACID} + \text{CARBONATE} \rightarrow \text{SALT} + \text{WATER} + \text{CO}_2 (g)
\]

**Example question:** Write an equation illustrating the reaction between solid calcium carbonate and hydrochloric acid solution*.

*See the ‘Aqueous Reactions’ notes for details on a slides and ladders exercise using this equation

**Details:** ‘Acid plus carbonate’ reactions are really a sequential double replacement / decomposition series:

**Task:** Balance and then combine the double replacement and decomposition reactions featured in the above example:

\[
\_ \text{HCl (aq)} + \_ \text{CaCO}_3 (s) \rightarrow \_ \text{CaCl}_2 (aq) + \_ \text{H}_2\text{CO}_3 (aq)
\]

\[
\_ \text{H}_2\text{CO}_3 (aq) \rightarrow \_ \text{H}_2\text{O (l)} + \_ \text{CO}_2 (g)
\]

\[
+ \_ \text{H}_2\text{O (l)} + \_ \text{CO}_2 (g)
\]
Background

‘REDOX’ reactions are chemical processes in which REDuction and OXidation simultaneously occur

Oxidation is Loss of electrons. An element or compound that loses electron(s) during a chemical process is said to be OXIDIZED

Reduction is Gain of electrons. An element or compound that gains electron(s) during a chemical process is said to be REDUCED

TRICK: Remembering the difference between oxidation and reduction is easy, just remember….

OIL RIG of electrons

Oxidation Is Loss, Reduction Is Gain of electrons
Overview Example of a Simple REDOX (combustion) reaction:

Discussion: Which chemical species has lost electrons during this process (i.e. been oxidized)? Which has gained electrons (been reduced)? How can you figure this out?

\[
2 \text{Mg(s)} + \text{O}_2(g) \rightarrow 2 \text{MgO(s)}
\]

Oxidized SPECIES = Oxidizing AGENT =

Reduced SPECIES = Reducing AGENT =

Notes
Oxidation Numbers

Oxidation Number or Oxidation State – the ‘charge’ on an atom in a substance as if it were a monatomic ion

A change in oxidation state during a chemical process indicates that a specie has either been oxidized (number goes up), or reduced (number goes down). Recall the previous MgO example.

Rules for assigning oxidation numbers

1. For materials that form atomic ions, the oxidation state is the same as the ‘regular’ ionic charge

Task: State the oxidation state of the following:

<table>
<thead>
<tr>
<th>Na in NaCl</th>
<th>Cl in AlCl$_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg in MgCl$_2$</td>
<td>Fe in Fe$_2$O$_3$</td>
</tr>
</tbody>
</table>

Since you know the charge of a great many atomic ions, you also know their oxidation states, i.e.:

- **Group I** = I (Li$^+$, Na$^+$…)
- **Group II** = II (Mg$^{2+}$, Ca$^{2+}$…)
- **Group VII** = -I (F$^-$, Cl$^-$…)
- **Group VI** = -II (O$^{2-}$, S$^{2-}$…)

Note: Oxidation states / numbers are expressed with Roman Numerals (this differentiates them from ‘pure’ ionic charges)
2. For ANY elemental atom, its oxidation state is ZERO. Why?

E.g. Elemental chlorine, \( \text{Cl}_2 \) ( \( \text{Cl} \leftrightarrow \text{Cl} \) )

Hint: Recall the trend in Electronegativity values (see appendix)

\[
\begin{array}{l}
\text{Any atom bonded to other identical atom(s) must have an oxidation state (oxidation number) of ZERO.}
\\
\text{ALL ELEMENTS must by definition posses zero oxidation states.}
\end{array}
\]

Examples: Any diatomic element (\( \text{O}_2, \text{F}_2 \)), any metallic element (\( \text{Pb}(s), \text{Al}(s) \)) etc.

3. All other atoms’ oxidation states must be determined mathematically using the ‘Sum of Oxidation States’ Rule:

\[
\begin{array}{l}
\text{For molecules: The sum of the molecule’s component atoms individual oxidation numbers} = \text{ZERO}
\end{array}
\]

Example: Nitric acid, \( \text{HNO}_3 \)
For polyatomic ions (‘charged molecules’): The sum of the polyatomic ion’s component atoms individual oxidation numbers = overall ionic charge

Example: The nitrate ion, NO$_3^-$

Exceptions:

Oxygen always has a -II oxidation state, except when bonded to either fluorine or itself. Why?

Hint: Think of the periodic trend in electronegativity (appendix).

Examples:

H$_2$O$_2$  OF$_2$
F always has a –I oxidation state, except when bonded to itself. Other Halogens (Cl, Br, I) are also –I, except when bonded to F or O. Why?

Examples:

\[
\text{ClO}_2 \quad \text{ClO}_3^- \\
\]

More Examples: Calculate the oxidation state of:

S in SO\textsubscript{3} \quad \text{Xe in XeF}_6

S in SO\textsubscript{4}^{2-} \quad \text{Cr in CrO}_4^-

Na in NaH (Hint: think electronegativity) \quad \text{N in Mg}_3\text{N}_2
Understanding REDOX reactions

Since ‘battery’ and single replacement reactions are examples of REDOX processes, oxidation numbers can be assigned. Comparing oxidation numbers of reactants and products shows which species have been oxidized and which reduced.

Example: Zn/Fe\(^{2+}\) ‘battery’ style REDOX reaction

\[
\text{Zn}(s) + \text{Fe}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Fe}(s)
\]

\[
\begin{array}{c}
0 \quad +2 \\
\text{Oxidation} \quad \text{Reduction}
\end{array}
\]

Oxidizing agent: Fe\(^{2+}\)
Reducing agent: Zn

Recall: For single replacement (‘prom’) reactions, the more reactive specie (A) undergoes oxidation when replacing the less reactive one (B), which is reduced.

\[\text{A} + \text{BX} \rightarrow \text{AX} + \text{B}\]

Example:

\[
\text{Mg}(s) + 2 \text{H}_2\text{O}(l) \rightarrow \text{Mg(OH)}_2(aq) + \text{H}_2(g)
\]

Oxidation states:

\[
\begin{array}{c}
0 \quad +1 \quad -2 \\
\text{Oxidation} \quad \text{Reduction}
\end{array}
\]
Tasks:

Assign oxidation numbers to the reactants and products in the following examples. State which specie is oxidized and which is reduced in each case.

\[ \text{Zn (s) } + \text{ Cu}^{2+} (\text{aq}) \rightarrow \text{ Cu (s) } + \text{ Zn}^{2+} \]

\[ \text{Ca (s) } + 2 \text{ HCl (aq) } \rightarrow \text{ CaCl}_2 + \text{ H}_2 (\text{g}) \]

\[ \text{Fe (s) } + \text{ Ni(NO}_3)_2 (\text{aq}) \rightarrow \text{ Ni (s) } + \text{ Fe(NO}_3)_2 (\text{aq}) \]

\[ \text{SO}_3 (\text{g}) + \text{ H}_2\text{O (l)} \rightarrow \text{ H}_2\text{SO}_4(\text{aq}) \]
Appendix: Electronegativity values
Gases

Reading: Ch 5, sections 1-7  
Homework: Chapter 5: 29, 33, 35, 37, 39, 41*, 43*, 45*, 47, 53, 55, 57, 61, 67, 69*, 75*, 77*  
* = ‘important’ homework question

Background

Discussion: What do we already know about gases? Many of the concepts to be covered are based on the interpretation of ‘everyday’ experiences.

Macroscopic View:  
Microscopic View:

1 atm = 760 mmHg (Torr) = 101 kPa

The behavior of a gas can be explained in terms of:

1. Its VOLUME (in liters)
2. Its PRESSURE (most often in atm.)
3. Its TEMPERATURE (in Kelvin)
4. The number of MOLES (n) of gas present

Understanding gases is all about understanding the relationships between the VOLUME, PRESSURE, TEMPERATURE and # MOLES (n) of a gas sample.
The Gas Laws

**Avogadro’s Law:** Relationship between \# **MOLES** of gas and **VOLUME** (P and T fixed)

*Everyday observation:* What takes up more space, 1 mole of gas or 2 moles of gas (assume constant temp. and pressure)?

\[ V \propto n \text{ (# moles)} \]

At standard temperature (273 K) and pressure (1.0 atm.), 1 mole of ANY gas occupies 22.4 L.

1 mole = 22.4 L at STP
Examples:

1. What volume does 3.5 moles of H$_2$ (g) occupy at STP?

2. How many atoms of He are there in a 5.0 L party balloon at STP?
Boyles Law: Relationship between PRESSURE and VOLUME (n and T fixed)

Everyday observation: If you compress (‘squeeze’) the container a gas is in, does the gas pressure increase or decrease (assume fixed T and n).

\[ P \propto \frac{1}{V} \quad \text{or} \quad PV = \text{constant (table)} \]

Example: A sample of gas has a pressure of 2.0 atm. If the volume of the container the gas sample is in is decreased by a factor of 10, what is the new pressure of the gas inside? Assume T and n constant.
Charles’ Law: Relationship between:
TEMPERATURE and VOLUME (n and P fixed) or
TEMPERATURE and PRESSURE (n and V fixed)*

Everyday observation: Why does a hot air balloon rise? What time of day is best for ballooning?

**Example:** A sample of gas occupies 12 L at 31°C. What volume does it occupy at 62°C? Assume P and n are constant.
The Ideal gas Law

The three gas law equations can be combined to make a new equation (the IDEAL GAS LAW) that can be used to solve ANY ‘static’ gas problem

Derivation of the ideal gas law (see appendix)

\[ PV = nRT \]

Where: \( R = 0.08206 \text{ L atm/mol K if the atm. pressure unit is used} \)
\[ \text{or} \quad R = 8.314 /\text{mol K if the Pa or N/m}^2 \text{ pressure unit is used} \]

Example: Calcium carbonate decomposes when heated to give solid calcium oxide and carbon dioxide gas. If 250 mL of \( \text{CO}_2 \) (g) is collected at 31°C and 1.3 atm pressure, then how many moles of \( \text{CO}_2 \) (g) is collected?

Write down what you are given, then solve for the unknown quantity in ‘static’ problems

\[ P = \]
\[ V = \]
\[ n = \]
\[ R = \]
\[ T = \]
Dynamic Problems: changing P, V or T for a fixed number of moles of gas

Derivation of the ‘Dynamic’ (before and after) gas law

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]
(before) \hspace{1cm} (after)

Example: The pressure inside an aerosol can is 1.5 atm. at 25°C. What would the pressure inside the can be at 450°C? *Is there a moral to the story?*

As with the ‘static’ problems, write down what you are given *then* solve for the unknown quantity in ‘dynamic’ problems

\[
P_1 = \hspace{1cm} P_2 = \\
V_1 = \hspace{1cm} V_2 = \\
T_1 = \hspace{1cm} T_2 =
\]
(before) \hspace{1cm} (after)
Measuring Atmospheric and/or Gas Pressure

Statement: The mmHg pressure unit sounds ‘all wrong’ – isn’t mm a length unit?

Reply: Yes, mmHg is a length unit, but it is related to the pressure of gas as measured by the height of a mercury barometer or manometer column (where 1 atm = 760mmHg)

Discussion: How does a mercury barometer work?

Discussion: Why is mercury used in barometers and manometers – why not water or some other ‘safe’ liquid?? Hint: what physical property is greatly different for Hg compared to other liquids?

The height of a column of liquid used to measure pressure and it’s density are related through the equation:

\[ P = gdh \]

Where \( P \) = pressure, \( g \) = gravitational constant (9.81ms\(^{-2}\)) and \( d \) = density of the liquid in the column.
Interesting question: Suppose water was used in place of mercury in all manometers and barometers. What then would be standard atmospheric pressure (760mmHg) be if measured in mmH$_2$O? The density of water is 1.000 g/mL, the density of Hg = 13.596 g/mL).

\[ \text{Ans: 10.33 m} \]

**The mercury manometer**

‘Before’ The pressure of the gas in the bulb and the atmospheric pressure are equal:

\[ P_{\text{gas}} = P_{\text{atm}} = 760 \text{ mmHg} \]

Therefore, the Hg in the U-tube is at the same level on the ‘gas side’ and the ‘atmosphere side’.

‘Change’: The pressure of the gas in the bulb is increased (how?). The extra pressure forces the Hg column down on the ‘gas side’ and ‘up’ on the ‘atmosphere side’ (like drinking a soda?!).

The *extra* pressure exerted by the gas is measured as a difference in height (\( \Delta h = 50 \text{ mm} \)) between the respective Hg levels

**After:** The new pressure of the gas is now its original pressure *plus* the increase in pressure:

\[ P_{\text{gas}} = 760 \text{ mmHg} + 50 \text{ mmHg} = 810 \text{ mmHg} \]
Applications of the Ideal Gas Equation

Determining the molar mass ($\mathcal{M}$) of a gas or vapor

Molar mass ($\mathcal{M}$) is defined as grams of chemical per mole of chemical. We can use the ideal gas law ($PV=nRT$) and the expression for the density of a gas ($D=M/V$) to find its molar mass.

Derivation:

1. Rearrange the ideal gas equation to make moles ($n$) the subject.

2. Rearrange the density equation to make mass the subject.

3. Divide equation 2 (grams) by equation 1 (moles) to find Molar mass.

Note: Equation 3 can be rearranged to make gas density the subject. This form of the equation is also used in problems (see appendix)
Example: What is the density of CCl₄(g) at 714 Torr and 125°C?

Answer: 4.43g/L (4.43 x10⁻³ g/mL)

Partial Pressure

Discussion: ‘Dry’ air is composed of 78% N₂, 21 % O₂ and 1% Ar. If the total atmospheric pressure exerted by the ‘dry’ air is 760 mmHg, then what fraction of this pressure is due to each component?

pN₂ =

pO₂ =

pAr =

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The total pressure \((P_{\text{total}})\) exerted by a gas mixture = sum of each component’s partial pressures \((p_1 + p_2 + p_3 \ldots\ldots)\):

\[
P_{\text{total}} = p_1 + p_2 + p_3 \ldots\ldots
\]

For air: \(P_{\text{air}} = p_{N_2} + p_{O_2} + p_{Ar} \)

i.e \(760\text{mmHg} = 593\text{mmHg} + 160\text{mmHg} + 7\text{mmHg}\)

Collecting gases ‘over water’: In the lab, gaseous reaction products are most often collected ‘over water’ (recall your ‘Determination of R’ lab)

Because the gas being collected has been passed through water, it ‘drags’ some water vapor (‘humidity’) with it. For the collection of a gas over water:

\[
P_{\text{wet gas}} = p_{\text{dry gas}} + p_{\text{water vapor}}
\]

Just like on a humid summer’s day, the amount (partial pressure) of \(H_2O\) vapor (gas) in the air is related to the ambient temperature:

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
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<td>71.97</td>
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</tr>
<tr>
<td>50</td>
<td>92.6</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
If the partial pressure of any components of a gas mixture (including ‘wet’ gases) are known, this information may be used in ideal gas law (PV = nRT) problems. Mole fractions may also be determined.

Example: During an experiment, H\(_2\)(g) was collected ‘over water’. The wet gas recovered was found to have a pressure of 745 Torr at 25°C.

1. What was the partial pressure of the dry H\(_2\) gas?

2. If the container the gas was collected in had a volume of 1.50 L, then how many moles of dry gas were collected?

3. What is the mole fraction of H\(_2\) (g) in the wet gas? What is the mole fraction of H\(_2\)O (g)?

Since \( P_{\text{comp.}} \propto n_{\text{comp.}} \) for a mixture of gases (with V and T fixed), then:

\[
\text{Mole fraction}_{\text{comp.}} = \frac{n_{\text{comp.}}}{n_{\text{mixture}}} = \frac{P_{\text{comp.}}}{P_{\text{mixture}}}
\]
The following reaction between calcium hydride and water is used to inflate life rafts and weather balloons:

\[
\text{CaH}_2(s) + 2 \text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(\text{aq}) + 2 \text{H}_2(\text{g})
\]

If 47.0 grams of CaH\(_2\)(s) is completely reacted with an excess of water, then:

A. How many moles of hydrogen gas will be produced?

B. What volume (in L) would the hydrogen gas generated in part (a) occupy at 15.0°C and 725 Torr?
Appendix
**Ideal Gas Law**

\[ PV = nRT \]

- For constant \( n \) and \( T \):
  \[ V = \frac{nRT}{P} \]
  \[ V \propto \frac{1}{P} \]
  Boyle's Law

- For constant \( n \) and \( P \):
  \[ V = \frac{nRT}{P} \]
  \[ V \propto T \]
  Charles's Law

- For constant \( P \) and \( T \):
  \[ V = \frac{nRT}{P} \]
  \[ V \propto n \]
  Avogadro's Law
Density

\[ n \frac{M}{V} = \frac{PM}{RT} \]

Molar density  Molar mass

\[ d = \frac{PM}{RT} \]

\[ \text{moles liter} \times \text{grams mole} = \text{grams liter} \]

Molar density  Molar mass  Density in grams/liter
Thermochemistry

<table>
<thead>
<tr>
<th>Reading: Ch 6, sections 1 – 10</th>
<th>Homework: Chapter 6: 33, 45, 47*, 49, 57, 59*, 61, 63, 65, 67*, 69, 73, 75*, 77*, 81, 83, 87</th>
</tr>
</thead>
</table>

* = ‘important’ homework question

Forms of Energy and Their Conversion

Discussion: Everyday physical objects, such as a golf ball flying through the air), posses two basic forms of energy – what are they?

1. 

2. 

‘Stunt Fish’ Demo and Analogy to a Chemical Systems

When energy is transferred or converted to another form, it must eventually appear as work and/or heat (think about your car)
Recall: The potential energy (enthalpy, H) possessed by elements and compounds is ‘stored’ within their chemical bonds.

**Breaking a chemical bond requires energy (heat)**

&

**Making a chemical bond releases energy (heat)**

Each type of chemical bond has it’s own enthalpy (stored chemical potential energy, e.g. O-O = 138 kJ/mole). See Appendix and the ‘More Chemical Bonding’ note packet for details.

Mathematical Version – The First Law of Thermodynamics (see Appendix)

\[ \Delta E = q + w \]

Where:

- \( \Delta E \) = change in energy for the system (‘dropping the fish’)
- \( q \) = heat energy transferred
- \( w \) = work performed (PV work for chemical systems, like for the combustion of gasoline in a car engine)

For chemical systems, there is most often no gas evolved, so there is no ‘PV’ work. Thus…..

\[ \Delta E = -\Delta H = +q_p \] (first law of thermodynamics)

Where:

- \( \Delta H \) = change in ‘chemical potential energy’ (enthalpy) for the system
- \( q_p \) = heat energy transferred to or from the chemical system (at constant pressure – i.e. no PV work)

Note: sign convention – heat energy is given out (+q) when the internal chemical energy (enthalpy) of the chemical system is reduced (-\( \Delta H \)). This is what we know as an exothermic process – more on this later.
Exothermic and Endothermic Chemical Processes

Definitions: Before we go any further, it is important to know a few key definitions. Hint: think about the construction of the words themselves to determine their meaning.

**Thermochemistry:** Study of the quantity of heat absorbed or evolved during a chemical reaction

**Thermodynamics:** Study of the relationship between heat and the other forms of energy involved in a chemical (or physical) process

**Heat:** The energy that flows into or out of a system because of a difference in temperature between the thermodynamic system and its surroundings. See slide.

**Thermodynamic system (system):** The substance or mixture of substances under study in which (chemical) change occurs

**Surroundings:** Everything in the vicinity of the thermodynamic system. Heat is either lost or gained to or from the surroundings. See slide.
Enthalpies of Reaction

Recall: All chemical reactions either release (exothermic) or absorb (endothermic) heat energy when going from reactants → products. As we saw in earlier material, this information can be represented quantitatively with a thermochemical equation.

Example: \[ \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (g); \Delta H = -890.3 \text{ kJ} \]

One mole of CH\(_4\) (g) reacts with 2 moles of O\(_2\) (g) to produce chemical products and 890.3 kJ of heat.

Questions: How much heat would be evolved if:

1. 2 moles of CH\(_4\) (g) were combusted in XS oxygen gas?

2. 4 moles of CH\(_4\) (g) were combusted in XS oxygen gas?

**Enthalpy is an extensive property** – it depends on the amount of material involved. “Two logs thrown on the fire produces twice as much heat as one log”

Change in Enthalpy, ΔH

Discussion: What does ΔH really mean, in terms of what’s going on with the reactants and products in a chemical reaction?
The change in enthalpy ($\Delta H$) is simply a measure of how much ‘chemical potential energy’ has been either lost or gained by converting the bonds present in the reactants ($H_{\text{reactants}}$) into the bonds present in the products ($H_{\text{products}}$). Recall that this difference in energy is either lost or gained from the system as heat.

Mathematically

$$\Delta H = (H_{\text{products}} - H_{\text{reactants}})$$

Recall also that:

$$-\Delta H = +q$$

Since $q$ can be measured experimentally (recall your lab), details pertaining to the enthalpy of the reactants and products involved in a chemical reaction can be determined. This theory underpins all calorimetric investigations.

Calorimetry

The amount of heat energy transferred to or from any material or object (thermodynamic system) can be found if its HEAT CAPACITY ($C_p$), MASS ($g$) and observed TEMPERATURE CHANGE, $\Delta T$ ($^\circ C$ or K), it undergoes are known:

$$q = C_p \times m \times \Delta T$$

Where:

$q =$ heat energy transferred to or from the system

$C_p =$ specific heat capacity of the system – a constant with units of $\text{J/g}^\circ\text{C}$

$m =$ mass of system in grams

$\Delta T =$ change in temperature (in K or $^\circ\text{C}$) of the system.

Defined at $T_{\text{final}} - T_{\text{initial}}$
Since each material has its own specific heat capacity, similar math applies:

**Group work**

1. How much heat energy is needed to raise the temperature of 25 g water by 15°C?

2. How much heat energy is needed to raise the temperature of 25 g solid iron by 15°C?
3. How much heat energy would be needed to boil a 330 mL glass of water that is initially at room temperature (25 °C)? Density H₂O (l) = 1.00 g/mL

4. A blacksmith tosses a ‘red hot’ iron horseshoe weighing 0.569 kg into a 5.0L bucket of water. If the water in the bucket rises in temperature from 10°C to 20°C, what was the original temperature of the horseshoe? Assume heat is only exchanged between the horseshoe and the water and no heat is lost to the surroundings.
Application to Chemical Systems

Recall: In your recent lab, the change in enthalpy (-ΔH) for the reaction between magnesium and HCl = heat energy (+q) gained by the surrounding solution in the insulated cup.

i.e.

\[- \Delta H_{\text{reaction}} = +q_{\text{solution}}\]

Remember: q is measured in ________, the S.I. unit of energy, even though thermodynamic values are typically quoted in ___________

Since the solutions used in calorimetry experiments typically contain a relatively low concentration of products:

\[C_p(\text{solution}) \approx C_p(\text{water}) = 4.18 \text{ J/g}^\circ\text{C}\]

Worked Example: 33 mL of 1.2 M HCl (aq) is added to 42 mL of a solution containing an excess of NaOH (aq). If a temperature change of 25°C → 31.8 °C is observed, calculate ΔH for the reaction. Quote your answer in kJ/mole.
Measuring Heats of Reaction (ΔH)

Calorimetry (measuring heats of reaction, recall your lab) can simply be considered as the practical application of the 1st Law, such as shown immediately above, i.e.:

$$-\Delta H_{\text{(reaction)}} = +q_{\text{(solution)}}$$

Where:

$$q_{\text{(solution)}} = C_p \times m \times \Delta T$$

The apparatus most often used for calorimetric measurements is a calorimeter – in your lab this was simply polystyrene coffee cup (see slide).

Calorimeter: “Device used to measure the heat absorbed or evolved during chemical or physical change”

Discussion: Can any container be used as a calorimeter? What is required of a vessel that is to be used as a calorimeter?

Types of calorimeter (see Appendix)

‘Coffee cup’ calorimeter (solutions)

‘Bomb’ calorimeter (gasses / solids)
Note: For a bomb calorimeter, the ‘math’ is slightly different, as the apparatus as a whole absorbs the heat evolved from the reaction. Thus:

\[ q_{\text{solution}} = C_{\text{cal}} \Delta T \]

Where: \( C_{\text{cal}} \) is the heat capacity (in J/°C or kJ/°C) of the calorimeter itself

Hess’ Law – Theoretical Determination of Heats of Reaction (\( \Delta H \))

Overview: We will learn and implement some familiar math to find \( \Delta H \) for reactions of interest – this is Hess’ Law.

Important fact: Enthalpy is an example of a State Function (see appendix / slide). This fact makes the math possible.

**Enthalpy as a state function:** ‘It doesn’t matter how you get there - it takes the same amount of energy (\( \Delta H \)) whatever route is taken’

Example: consider the following number lines, which represent enthalpy changes during chemical reaction(s)

1.

2.

**In each case \( \Delta H \) is the same, irrespective of the route taken. This is of use, since we can find an unknown \( \Delta H_{\text{rxn}} \) from established values – this is Hess’s law of Heat summation**
Hess’s law of Heat summation: *For a chemical equation that can be written as the sum of two or more steps, the enthalpy change for the overall reaction equals the sum of the enthalpy changes for the individual steps*

‘Faulty water heater’ example: Find $\Delta H_{\text{rxn}}$ for the following:

$$2 \text{C (graphite) + O}_2 (g) \rightarrow 2 \text{CO (g)}; \Delta H_{\text{rxn}} = ?$$

Given:

1. $2 \text{C (graphite) + 2 O}_2 (g) \rightarrow 2 \text{CO}_2 (g); \Delta H_1 = -787 \text{ kJ}$
2. $2 \text{CO}_2 (g) \rightarrow 2 \text{CO (g) + O}_2 (g); \Delta H_2 = +566 \text{ kJ}$

Since the desired reaction is the sum of reactions 1 and 2, then

$$\Delta H_{\text{rxn}} = \Delta H_1 + \Delta H_2$$

Manipulating Thermochemical equations

Sometimes just adding known thermochemical equations does not give the desired unknown thermochemical equation. The known thermochemical equations can be ‘worked over’ by following these rules:

- **Add Reactions**: Add $\Delta H$s
- **Subtract Reactions**: Subtract $\Delta H$s
- **Multiply Reaction**: Multiply $\Delta H$
- **Reverse Reaction**: Multiply $\Delta H \times (-1)$
**Worked Example:** Find $\Delta H_{\text{rxn}}$ for the following:

$$2 \text{S (s)} + 3 \text{O}_2 (g) \rightarrow 2 \text{SO}_3 (g); \quad \Delta H_{\text{rxn}} = ?$$

**Given:**

$$\text{S (s)} + \text{O}_2 (g) \rightarrow \text{SO}_2 (g); \quad \Delta H_1 = -297 \text{ kJ}$$

$$2 \text{SO}_3 (g) \rightarrow 2 \text{SO}_2 (g) + \text{O}_2 (g); \quad \Delta H_2 = +198 \text{ kJ}$$

---

**~Always follow this procedure:**

1. Arrange the known reaction(s) so reactant(s) and product(s) common to the unknown reaction appear on the ‘correct’ sides.

2. Multiply known reaction(s) so similar amounts of reactants and products, common to the unknown reaction, appear on the ‘correct’ sides.

3. Add known reactions – their product should similar to the unknown
Task: Calculate $\Delta H_{\text{rxn}}$ for:

$$4 \text{ Al (s)} + 3 \text{ MnO}_2 (s) \rightarrow 2 \text{ Al}_2\text{O}_3 (s) + 3 \text{ Mn (s)} ; \quad \Delta H_{\text{rxn}} = ?$$

Given:

$$2 \text{ Al (s)} + \frac{3}{2} \text{ O}_2 (g) \rightarrow \text{ Al}_2\text{O}_3 (s) ; \quad \Delta H_1 = -1676 \text{ kJ}$$

$$\text{ Mn (s)} + \text{ O}_2 (g) \rightarrow \text{ MnO}_2 (s) \quad \Delta H_2 = -521 \text{ kJ}$$
Idea / Analogy: Establish a ‘spelling bee champion vocabulary’ of known enthalpies (standard heats of formation, $\Delta H_f$) from which any ‘sentence’ (reaction) can be constructed.

Analogy: The following two short ‘sentences’ (reactions) can be combined to give the required ‘longer sentence’ (equation):

$$\begin{align*}
\text{The cat sat} \\
\text{on the mat} & + \\
\text{The cat sat on the mat}
\end{align*}$$

Recall: This is how our previous Hess’ Law examples have worked to this point. However, the two small ‘sentences’ can only be combined to make the final The cat sat on the mat sentence. This is somewhat limiting, as other new sentences cannot be constructed.

However, having a large vocabulary (knowing lots of words) means virtually any new sentence can be constructed, for example Mat sat on the cat!

Standard heats of formation are the chemical equivalent of words in our analogy

**Definition of Standard Heats (Enthalpies) of Formation ($\Delta H_f$)**

The enthalpy change (kJ/mol) for the formation of one mole of a substance in its standard state (i.e. its physical state at 1.00 atm, 25°C) from its elements in their standard states.

**Example:** $\text{H}_2 (\text{g}) + \frac{1}{2} \text{O}_2 (\text{g}) \rightarrow \text{H}_2\text{O} (\text{l}) \ ; \Delta H_f = -285.8 \text{ kJ}$

**Discussion:** Some heats of formation have zero values. Why is this? See Appendices here & you’re text for a full list

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Recall that enthalpy is a state function, so the enthalpy change for reactants → component elements (-ΔH_f reactants) → products (ΔH_f products) can be determined for any reactants → products (ΔH_rxn) reaction. 

(ΔH_rxn) values can be determined from ΔH_f values in any one of two ways – graphically or via the use of a formula.

Simple worked example: What is the heat of vaporization, ΔH_vap (the energy needed to convert 1 mole of liquid substance to one mole of gaseous substance at its boiling point) for:

CS_2(l) → CS_2(g); ΔH_vap = ?

Where: ΔH_f CS_2(l) = +87.9 kJ and ΔH_f CS_2(g) = +117 kJ

Graphical method:

Formula method:

From inspection above, it can be seen that, essentially:

ΔH_rxn = - ΔH_f (reactants) + ΔH_f (products)

Taking into account possible multiple reactant / product species and their respective stoichiometric constants, we arrive at:

ΔH_rxn = ΣnΔH_f (products) - ΣmΔH_f (reactants)
Task: calculate the heat of vaporization for water. What is the heat of vaporization for 2 moles of water?

Where: $\Delta H_f \text{H}_2\text{O}(l) = -285.8 \text{ kJ}$ and $\Delta H_f \text{H}_2\text{O (g)} = -241.8 \text{ kJ}$

ANS = 44 kJ
Harder worked example: calculate $\Delta H_{\text{rxn}}$ for:

$$4 \text{NH}_3 (g) + 5 \text{O}_2 (g) \rightarrow 4 \text{NO} (g) + 6 \text{H}_2\text{O} (g)$$

Given:

- $\Delta H_f \text{NH}_3 (g) = -45.9 \text{ kJ}$
- $\Delta H_f \text{NO} (g) = +90.3 \text{ kJ}$
- $\Delta H_f \text{H}_2\text{O} (g) = -241.8 \text{ kJ}$
Group Task: calculate $\Delta H_{\text{rxn}}$ for:

$$2 \text{ PbS (s) } + \text{ 3 O}_2 (\text{g}) \rightarrow 2 \text{ SO}_2 (\text{g}) + 2 \text{ PbO (s)}$$

Given:

$\Delta H_f \text{ PbS (s) } = -98.3 \text{ kJ}$

$\Delta H_f \text{ SO}_2 (\text{g}) = -296.8 \text{ kJ}$

$\Delta H_f \text{ PbO (s) } = -219.0 \text{ kJ}$

ANS = $-835.6 \text{ kJ}$
Appendix: The Strength of Covalent Bonds (Bond Enthalpies)

Background: Enthalpy the energy (in kJ) required to break one mole of a specified type of bond. Units are kJ/mole. For Cl₂:

\[ \text{Cl} - \text{Cl} \rightarrow \text{Cl} + \text{Cl} \; \Delta H = 242 \text{ kJ/mol} \]

i.e. it takes 242 kJ of energy to break one mole of Cl-Cl bonds

### Some Average Single- and Multiple-Bond Energies*

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<td>I  201</td>
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<tr>
<td>O≡O (in O₂) 498</td>
<td>C≡O 1075</td>
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</table>

*In kilojoules per mole.
**A State Function**

Change in altitude depends only on the difference between the initial and final values, not on the path taken.

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<table>
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<th>Formula</th>
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<th>$\Delta H_f$ (kJ/mol)</th>
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<td>−273.3</td>
<td>Na$_2$CO$_3$(s)</td>
<td>−1130.7</td>
</tr>
<tr>
<td>CH$_4$(g)</td>
<td>−74.6</td>
<td>Hydrogen</td>
<td></td>
<td>NaHCO$_3$(s)</td>
<td>−950.8</td>
</tr>
<tr>
<td>CH$_2$OH(l)</td>
<td>−238.6</td>
<td>H(g)</td>
<td>218.0</td>
<td>Sulfur</td>
<td></td>
</tr>
<tr>
<td>C$_2$H$_5$(g)</td>
<td>227.4</td>
<td>H$_2$(g)</td>
<td>0</td>
<td>S$_8$(s, rhombic)</td>
<td>0</td>
</tr>
<tr>
<td>C$_2$H$_5$(g)</td>
<td>52.4</td>
<td>Nitrogen</td>
<td></td>
<td>S$_8$(s, monoclinic)</td>
<td>0.3</td>
</tr>
<tr>
<td>C$_2$H$_6$(g)</td>
<td>−84.68</td>
<td>N$_2$(g)</td>
<td>0</td>
<td>SO$_2$(g)</td>
<td>−296.8</td>
</tr>
<tr>
<td>C$_2$H$_5$OH(l)</td>
<td>−277.6</td>
<td>NH$_3$(g)</td>
<td>−45.9</td>
<td>SO$_3$(g)</td>
<td>−395.7</td>
</tr>
<tr>
<td>C$_2$H$_6$(g)</td>
<td>−103.85</td>
<td>NH$_2$NO$_3$(s)</td>
<td>−365.6</td>
<td>H$_2$SO$_4$(l)</td>
<td>−814.0</td>
</tr>
<tr>
<td>C$_3$H$_6$O$_3$(lac.)</td>
<td>−248.4</td>
<td>NO(g)</td>
<td>91.3</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>N$_2$O(g)</td>
<td>81.6</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
The Coffee-Cup Calorimeter

- Thermometer
- Glass stirrer
- Cork lid (loose fitting)
- Two nested Styrofoam® cups containing reactants in solution

Copyright © 2008 Pearson Prentice Hall, Inc.
The Bomb Calorimeter

\[ q = C \times \Delta T \]

Heat capacity

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<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat Capacity, $C_s$ (J/g • °C$^*$)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Elements</strong></td>
<td></td>
</tr>
<tr>
<td>Lead</td>
<td>0.128</td>
</tr>
<tr>
<td>Gold</td>
<td>0.128</td>
</tr>
<tr>
<td>Silver</td>
<td>0.235</td>
</tr>
<tr>
<td>Copper</td>
<td>0.385</td>
</tr>
<tr>
<td>Iron</td>
<td>0.449</td>
</tr>
<tr>
<td>Aluminum</td>
<td>0.903</td>
</tr>
<tr>
<td><strong>Compounds</strong></td>
<td></td>
</tr>
<tr>
<td>Ethanol</td>
<td>2.42</td>
</tr>
<tr>
<td>Water</td>
<td>4.18</td>
</tr>
<tr>
<td><strong>Materials</strong></td>
<td></td>
</tr>
<tr>
<td>Glass (Pyrex)</td>
<td>0.75</td>
</tr>
<tr>
<td>Granite</td>
<td>0.79</td>
</tr>
<tr>
<td>Sand</td>
<td>0.84</td>
</tr>
</tbody>
</table>

*At 298 K.

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Modern Atomic Theory

<table>
<thead>
<tr>
<th>Reading:</th>
<th>Chapter 9: sections 1-4</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Ch. 7, sections 5-6 (lec)</td>
</tr>
<tr>
<td></td>
<td>Ch. 7, sections 1-3 (lab)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Homework:</th>
<th>Chapter 9: 37*, 39*, 41</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Chapter 7: 59, 61*, 63, 65 (lec.)</td>
</tr>
<tr>
<td></td>
<td>Chapter 7: 39, 41, 43, 47 (lab)</td>
</tr>
</tbody>
</table>

* = ‘important’ homework question

Note: Much of the early reading from Ch9. pertains to material covered in the next ‘applications’ packet.

The Electronic Structure of Atoms – Empirical ‘Particle’ Models

Recall: What called Mendeleev to stack certain elements in ‘Family’ groups within the periodic table?

Macroscopic properties: Elements in the same group (column) of the periodic table have similar chemical and physical properties.

e.g. All the group II elements (the alkali earths) are all metallic, form alkali solutions when mixed with water and form +2 charge cations when ionized.
Why is this? What is the underlying microscopic ‘answer’ that explains these facts?

**Microscopic properties:** Elements in the same group (column) of the periodic table have similar outer (valence) electronic configurations.

<table>
<thead>
<tr>
<th>Key</th>
<th>The loss /gain (ionic bonding), or the sharing (covalent bonding), of an atom’s outer most electrons IS chemistry.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Chemistry is ‘all about’ making new materials (as described by a chemical reaction’s equation) – to do this new bonds are made (in the products) and old bonds are broken (in the reactants). Bond making and breaking ONLY involves the outer valence electrons of atoms.</td>
</tr>
</tbody>
</table>

Empirical Electronic (‘Dot’) configurations of the first 12 elements (*leave space for more or write on the rear of this sheet!*)}
Seven Key facts

1.

2.

3.

4.

5.

6.

7.
Important definitions:

Electron ‘dot’ symbol: Includes BOTH outer (valence) AND inner (core) electrons

Lewis symbol: Includes outer (valence) electrons ONLY

Task: Draw electron ‘dot’ and Lewis symbols for:

Si

Cl

P

Add these Dot diagrams to your periodic table of electronic structure. Complete the table for all atom types up to and including Ca
Task: Complete the following table:

<table>
<thead>
<tr>
<th>Atom</th>
<th>Group number</th>
<th>Number of valence electrons</th>
<th>Lewis Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>N</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>P</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>S</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Si</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Recall that the number of valence electrons an atom has is equal to its group number – this is why elements in the same group have similar chemical properties (similar valence configurations)
The Octet Rule (Full Valence Shell rule – more detail later!)

ATOMS WITH FULL OUTER (VALENCE) SHELLS ARE STABLE ⇒ Atoms will lose, gain or share electrons to have an inert gas (full valence shell) configuration.

THIS IS THE ‘DRIVING’ FORCE BEHIND ALL CHEMICAL PROCESSES (see slide).

Examples:

1. Ionic bonding – the formation of LiF

Recall: Ionic bonds form between atoms (metal and non-metal, which then become ions) with a large difference in electronegativity

Task: Write a similar reaction for the formation of MgO.

2. Covalent bonding – the formation of F₂(g)

Recall: Covalent bonds form between atoms (two non-metals) with little or no difference in electronegativity

Note: We will return to constructing molecular models from Lewis symbols later in the course.
The Electronic Structure of Atoms – True ‘Wave’ Models

Background: The turn of the 20th century was a golden era for the physical sciences. Einstein was working towards the theory of special relativity (and was receiving all the attention), while Erwin Schrödinger was quietly ‘turning the science on its head’ by redefining how scientists fundamentally think about matter, particularly ‘small’ particles, such as electrons.

Schrödinger’s great genius was devising a form of math that explained the behavior of electrons (or any other ‘small particle’) in terms of WAVE, rather than PARTICLE, math. This new math is called quantum mechanics, with the main equation used to solve such problems called the Schrödinger equation. The affects of this ‘quiet revolution’ totally changed how we now view the structure of the atom.

Instead of electrons ‘orbiting’ a nucleus (PARTICLE MODEL), they are now known to exist as waves in diffuse clouds around the nucleus (WAVE MODEL).

Got proof? Yes!
The Schrödinger Equation

\[ \frac{-\hbar^2}{2m} \nabla^2 \psi + V \psi = E \psi \]

The Schrödinger equation, as seen on his and hers ‘nerd’ T-shirts at Star Trek Conventions, Physics conferences etc. (left and above). Hard core ‘Trekies’ (right)

The Power of the Schrödinger Equation

The Schrödinger equation supplies information pertinent to an electron(s)’ energy (E, V) and location (V) within ANY atom or molecule in terms of the behavior of electron ‘clouds’ or waves (Ψ).

Solving the Schrödinger equation for valence electrons tells you everything about their interactions with one another

**THIS IS CHEMISTRY**

We will use the RESULTS* from such calculations in order to better understand atomic structure (i.e. where electrons are located in atoms)

*ask me to tell you a scary story from grad. school!
Current Applications: The mathematical resources required to solve the Schrödinger equation are immense – until several years ago entire main frame computers were required to run for several days in order to ‘map’ a simple molecule. With the advent of modern high powered computers, such calculations can now be conducted much more quickly – chemical processes (such as drug synthesis) are now most often modeled using a computer before performing an experiment!

![An electron density map of a complex organic molecule – created on a computer by solving Schrödinger’s equation](image)

An overview comparison of ‘wave’ and ‘particle’ models

Background: the empirical Lewis and ‘dot’ models work (to a point) because they in many ways mimic ‘how it really is’ in terms of the electronic structure of the atom.

Objective: We will develop an understanding of the ‘true’ (wave) models of the atom by building on our knowledge of the more ‘common sense’ empirical (Lewis and ‘dot’) models.
Task: In your own words, compare and contrast the generic ‘particle’ and ‘wave’ models of an atom’s electrons, as shown below (see slide also)

We will return to comparisons like this periodically through the handout

Notes:
Quantum Numbers and Atomic Orbitals

The results of the Schrödinger equation (info. regarding electron position and energy in an atom) are expressed in terms of quantum numbers.

Each electron in an atom has a unique set of quantum numbers – they may be considered its ‘atomic address’ (the orbital it resides in)

There are five quantum numbers: n, l, ml and ms

Definition: An orbital is a region of space around the nucleus (with specific shape and direction) where (up to two) electrons can exist. When occupied by electron(s), orbitals become ‘electron clouds’. In the above example (H), the orbital is spherical in shape and contains 1 electron.

Analogy: Think of the nucleus as a sports stadium surrounded by numerous empty parking spaces (orbitals). Cars (electrons) fill some of these spaces (make electron clouds). The number of electrons attracted (cars in the lot) depends on the atom (1H, 2He etc.) The address of each ‘space’ (orbital), and therefore the car (electron) parked within it, is defined by a unique set of quantum numbers – discussed below.

Note: Every atom is assumed to have the same ‘infinitely’ large parking lot (number of available orbitals). ‘How many electrons’ and ‘where they go’ are what we really need to know for each atom – this is next!
Overview of an Orbital’s Quantum Numbers and their Meaning:

<table>
<thead>
<tr>
<th>Quantum number</th>
<th>Atomic Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>n</td>
<td>Average <strong>distance</strong> of orbital from nucleus</td>
</tr>
<tr>
<td>l</td>
<td><strong>Shape</strong> of orbital (spherical or ‘dumbbell’, see below)</td>
</tr>
<tr>
<td>m_l</td>
<td><strong>Direction</strong> of orbital in space (oriented in x, y or z direction)</td>
</tr>
<tr>
<td>m_s</td>
<td>‘<strong>Spin</strong>’ of electron(s) (↓ or ↑) residing within the orbital</td>
</tr>
</tbody>
</table>

Details regarding Orbitals and the Quantum numbers that describe them

Simple ideas (same as for the empirical models):

1. Electrons (-ve) are attracted to the nucleus (+ve).

2. The number of electrons attracted ≡ number of protons in the nucleus (atomic number, Z)

3. As with Dodger stadium, there is more space (available parking) further from the nucleus (stadium). **Diagram:**
Specific details (some analogies with classic ‘particle’ models)

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Associated Quantum number</th>
<th>Range of allowed values</th>
<th>Details</th>
</tr>
</thead>
<tbody>
<tr>
<td>‘Distance’</td>
<td>“Principle quantum number” ( n )</td>
<td>( n = 1, 2, 3… )</td>
<td>( n \propto ) distance orbital is from the nucleus ( n = 1 ) is the closest energy level Analogous to the empirical ‘dot’ model. Highest ( n ) = element’s row in p. table</td>
</tr>
</tbody>
</table>

Example (recall your Atomic Fingerprint lab):

As with the empirical ‘dot’ models, an electron’s distance from the nucleus relates to which ‘orbit’ (particle model) or energy level (wave model) it is in.

\( \Rightarrow \) The valence ‘orbit’ of Na (row 3 p. table) has \( n = 3 \). The core (full) levels are \( n=2 \) and \( n = 1 \).

<table>
<thead>
<tr>
<th>‘Shape’</th>
<th>“Angular momentum quantum number” ( l )</th>
<th>( l = 0 \rightarrow (n-1) )</th>
<th>I defines the shape of an orbital when ( l= 0 ), orbital is a spherical ‘s’ orbital when ( l= 1 ), orbital is a dumbbell ‘p’ orbital when ( l= 2 ), orbital is a fan/donut ‘d’ orbital Key: the value of ( n ) determines the range of ( l )</th>
</tr>
</thead>
</table>

Examples

A 1s (spherical) orbital A 2p (‘dumbbell’) orbital A 3d (‘fan’) orbital
| ‘Direction’ | “magnetic quantum number” $m_l$ | $m_l = 0 \rightarrow \pm 1$ | $m_l$ defines the direction of an orbital when $l=0$, $m_l=0$ (only one possible direction for ‘s’ orbitals) when $l=1$, $m_l=-1, 0$ or $+1$ (3 possible directions for ‘p’ orbitals) when $l=2$, $m_l=-2,-1,0,+1$ or $+2$ (5 possible directions for ‘d’ orbitals) **key:** the value of $n$ determines the range of $l$, which then determines the range of $m_l$ |

**Examples:**

**Only one possible direction for a spherical ‘s’ orbital, regardless of the value of $n$.**

**Three possible directions for ‘p’ orbitals ($p_x$, $p_y$, $p_z$) when $n=2$ or greater**

**Five possible directions for ‘d’ orbitals ($d_{xy}$, $d_{yz}$, $d_{xz}$, $d_{xy}$, $d_{x^2-y^2}$) when $n=3$ or greater**

---

**Pictorial review of the parking lot analogy**

**Dodger Stadium (park your car at…)**

**The Atom (park your electron at…)**
For the electrons themselves (the ‘cars’ that park in the ‘spaces’ detailed above)

| ‘Orientation’ | “magnetic spin quantum number” | $m_s$ | $m_s$ is an intrinsic property of the electron. Electrons are either ‘spin up’ (↑), $m_s = +\frac{1}{2}$ or spin down (↓), $m_s = -\frac{1}{2}$ | Analogy: It’s a bit like being either male (♂) or female (♀)!

**key:** ANY orbital can contain a MAX of 2 electrons – one ‘spin up’ (↑), $m_s = +\frac{1}{2}$ and one ‘spin down’ (↓), $m_s = -\frac{1}{2}$

Analogy: It’s a bit like having to get married before you move in together!

Example: Electrons either align with or against an external magnetic field, depending on whether they are respectively either ‘spin up’ ((↑), $m_s = +\frac{1}{2}$) or spin down ((↓), $m_s = -\frac{1}{2}$).

There are essentially equal numbers of each type of electron, which under certain circumstances can change from ‘spin up’ to ‘spin down’ (just like people!). However, there is no ‘gay marriage’ for electrons!

Typical ‘starter’ question: Sketch the shape and orientation of the following types of orbitals. See the appendix for more details.

\[ s \quad p_z \quad d_{xy} \]

America’s next top model is……

Schrödinger’s wave model is a more sophisticated version of the analogous Lewis style particle description.

In other words, both models work, but the wave model is a better description of how things really are in nature.
Putting it all together: a comparison of particle and wave models

Side by Side Comparison of the Two Types of Model

Particle model

Wave Model
Draw Lewis symbols and ‘dot’ structures and for the following:

<table>
<thead>
<tr>
<th>Lewis symbol</th>
<th>Dot structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon atom</td>
<td></td>
</tr>
<tr>
<td>Oxide anion</td>
<td></td>
</tr>
<tr>
<td>Sodium atom</td>
<td></td>
</tr>
<tr>
<td>Hydrogen atom</td>
<td></td>
</tr>
<tr>
<td>Magnesium ion</td>
<td></td>
</tr>
</tbody>
</table>
Appendix: s, p and d Orbital Plots

(a) 1s orbital

Density of dots proportional to probability density ($\psi^2$).

(b) Probability density ($\psi^2$)

Height of curve proportional to probability density ($\psi^2$).

The 2s and 3s Orbitals

2s ($n = 2, l = 0$)

3s ($n = 3, l = 0$)

Nodes

Total radial probability

Total radial probability

r (100 pm)

r (100 pm)
Modern Atomic Theory - Applications

**Reading:** Ch 7, sections 4-6  
Ch 8, sections 2-5

**Homework:** Chapter 7: 59†, 61†, 63†,  
Chapter 8: 43*, 45*, 47*, 49, 65

* = ‘important’ homework question. † = question repeated from previous ‘theory’ packet

**Overview:** The RESULTS from the Schrödinger Equation (the unique set \( n, l, m_l \) and \( m_s \) quantum numbers an electron in an atom possesses) allow for a variety of important determinations

Quantum number interdependency: ‘allowed’ quantum numbers for electrons in atoms.

The range of each ‘allowed’ set of quantum numbers for any electron is interdependent. There respective ranges are determined by the following key relationships:

- \( n = 1, 2, 3… \) The principle quantum number (‘shell’ or ‘layer’)
- \( l = 0 \rightarrow (n-1) \) ‘Shape’ quantum number (0 = ‘s’, 1 = ‘p’, 2 = ‘d’)
- \( m_l = 0 \rightarrow \pm l \) ‘# of Directions’ quantum number
- \( m_s = -\frac{1}{2} \text{ or } +\frac{1}{2} \) ‘Type’ of electron quantum number (\( \uparrow, +\frac{1}{2} \text{ or } \downarrow, -\frac{1}{2} \))

Example: Which of the following represents an impossible combination of \( n \) and \( l \)?

Recall the shorthand ‘code’ for assigning quantum numbers to specific orbitals (electron addresses):

- \( 1s \) (\( n = 1, l = 0 \)), \( 2p \) (\( n = 2, l = 1 \)), \( 2s \) (\( n = \_\_\_, l = \_\_\_ \)) OR

Recall that ‘more space = more shapes’ i.e.:

\[ \begin{array}{c|c|c}
\text{‘EZ’} & n = 1, s \text{ only} & n = 2, s, p \text{ only} & n = 3, s, p, d \text{ only} \\
\hline
\text{a. 1p} & \text{b. 4s} & \text{c. 2d} \\
\end{array} \]
As we will discover, quantum number interdependency describes the location and energy of electrons within atomic orbitals *and* underpins the design of the periodic table.

Quantum Mechanical Map of the Atom – Energy Level Diagrams

**Background:** An orbital ‘map’ of the atom (energy level diagram) can be constructed using quantum numbers.

**Task:** Understand the layout and construction of the energy level diagram (for orbitals up to n= 3) for hydrogen. See next page for details.
Recall what each quantum number represents, and how they are related to one another – this gives the distance, shape and direction of the allowed orbitals for each shell (n value). Include this information in the above energy level diagram. Recall that each orbital (box) can contain up to 2 e - we will use this fact later.

Energy Level Diagrams for Many Electron Atoms

And now we move the goal posts! For many electron atoms (everything but H), electron – electron interactions cause the ‘s’ orbital of each shell to contract, i.e. get closer to the nucleus. This is most pronounced for the 4s, which dips below the 3d (!) when the 3s and 3p levels are filled (see slide and appendix)

Task: Sketch the Energy Level diagram for a many electron atom. Compare and contrast this diagram with that for H.
Application 1: Electronic Configurations (‘filling the parking lot’)

The Energy level diagrams essentially give us a map of the atom’s orbitals (‘parking lot’). We can now fill these atomic orbitals with electrons (‘fill the parking spaces’) by applying some simple rules.

Rules for filling energy level diagrams

1. Number electrons ‘attracted’ \(\equiv\) Number protons in nucleus

2. Electrons fill lowest energy levels first (why?) – this is the Aufbau principle (‘building up’)

3. Electrons only pair up ‘when they have to’ (why?) – this is Hund’s rule

4. There can be no more than 2 electrons per orbital (why?) – this is the Pauli exclusion principle

Energy level diagram illustrating the order of filling orbitals

Summary of the order of filling orbitals
Task: Use the blanks provided to sketch energy level diagrams for the specified atoms

- **Nitrogen (7N)**
  - 1s
  - 2s
  - 3s
  - 3p

- **Fluorine (9F)**
  - 1s
  - 2s
  - 3s
  - 3p

- **Phosphorous (15P)**
  - 1s
  - 2s
  - 3s
  - 3p

- **Chlorine (17Cl)**
  - 1s
  - 2s
  - 3s
  - 3p

Question: Do you notice any similarities to the atoms’ respective ‘dot’ diagrams?
Orbital ‘box’ diagrams and electronic configurations

Once the order of filling orbitals is known (from energy level diagrams), this information can be presented in abbreviated form. *Orbital ‘box’ diagrams and electronic configurations* achieve this goal.

An *orbital box diagram* is simply a ‘linear version’ of the atom’s respective *energy level diagram*.

Example: Energy level and Orbital box diagrams for Oxygen

**Energy Level Diagram**

```
   Energy Level Diagram

   1s     2s     2p 3p
   1s  2s 2p 3p
     |     |     |
   1s  2s 2p 3p
   1s  2s 2p 3p
```

**‘Box’ Diagram**

```
O:     
1s     2s     2p
```

**Notes:**

More Examples:

**Li:**

```
Li: 1s 2s
    1 1
   1s 2s
```

**C:**

```
C: 1s 2s 2p
    1 1 1
   1s 2s 2p
```
An electronic configuration is a ‘shorthand’ version of the atom’s respective orbital box diagram. The number of electrons in each orbital or set of degenerate orbitals is represented by superscripts (these are not mathematical ‘powers’).

Examples

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Number of electrons</th>
<th>Electron configuration</th>
<th>Orbital diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>3</td>
<td>1s^22s^1</td>
<td><img src="image" alt="Orbital diagram for Li" /></td>
</tr>
<tr>
<td>Be</td>
<td>4</td>
<td>1s^22s^2</td>
<td><img src="image" alt="Orbital diagram for Be" /></td>
</tr>
<tr>
<td>B</td>
<td>5</td>
<td>1s^22s^2p^1</td>
<td><img src="image" alt="Orbital diagram for B" /></td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>1s^22s^2p^2</td>
<td><img src="image" alt="Orbital diagram for C" /></td>
</tr>
</tbody>
</table>

Task: Draw an Orbital box diagram and write an electronic configuration for Nitrogen

Orbital ‘Box’ Diagram

Electronic Configuration

N: 1s 2s 2p

Notes:
Condensed Electron Configurations

An atom’s core electrons have, by definition, an inert gas or complete shell configuration. The respective inert gas electron configuration of an atom’s core electrons can be substituted for the corresponding inert gas atomic symbol in a condensed electronic configuration.

Where: $[\text{He}] = 1s^2$, $[\text{Ne}] = 1s^2 2s^2 2p^6$, $[\text{Ar}] = 1s^2 2s^2 2p^6 3s^2 3p^6$

Worked Example: Silicon

Si: $1s^2 2s^2 2p^6 3s^2 3p^2$

Core electrons Valence electrons

Si: $[\ ] 3s^2 3p^2$

Core electrons Valence electrons

Worked Examples:

<table>
<thead>
<tr>
<th>Atom</th>
<th>Electron Configuration</th>
<th>Condensed Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>Li: $1s^2 2s^1$</td>
<td>Li: $[\text{He}] 2s^1$</td>
</tr>
<tr>
<td>Mg</td>
<td></td>
<td></td>
</tr>
<tr>
<td>K</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Details: Transition and other ‘heavy’ element and/or ion electronic configurations

The ‘d’ orbitals of any ‘post calcium’ element or ion, by definition, belong to the ‘core’ – WHY?

Answer:

Examples:

Ge

\[
1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2
\]

28 core electrons

4 valence electrons

Zn

[Ar] \(4s^2 3d^{10}\)

Zn\(^{2+}\)

[Ar] \(4s^0 3d^{10}\)

Task: Write a condensed electronic configuration for Bromine. How many valence electrons does this element have?
Application 2: Construction of the Periodic Table

The relationship between the arrangement of an atom’s total number of electrons and its position in the periodic table is well understood in terms of the empirical ‘battleship’ analogy (column 2 = 2 valence e, row 3 = 3rd shell etc.)

We can now build an analogous (non-empirical) model of the periodic table using quantum mechanical results in place of ‘dot’ diagrams.

‘Read across’ the period table while applying the filling of orbitals rules to obtain the electronic configuration of any element. This methodology is illustrated in the following version of the periodic table (see slide / appendix).
Worked Example: Chlorine

Electronic configuration of Chlorine

Cl: \( \text{or} \) Cl:

Determine the following atoms’ electronic configurations of the by ‘reading across’ the periodic table

Li: \( \text{or} \) Li:

Al: \( \text{or} \) Al:

Se: \( \text{or} \) Se:
(see next page for hint)

Xe: \( \text{or} \) Xe:
Review: ‘reading across’ to find the electronic configuration of Se

Note: The periodic table has a total of 7 rows. This means that the largest known atoms have a maximum of 7 shells. For the largest fully complete shells, l has a range of 0 → 4. The ‘new’ types of orbital now allowed (when l = 4) are \( f \) orbitals – filling these orbitals creates the Lanthanide (4\( f \)) and Actinide (5\( f \)) series, between the ‘s’ and ‘d’ blocks, in rows 6 and 7 respectively. Similar to the 3d elements (which appear in row 4), the ‘f’ block elements also appear in lower rows (e.g. 4\( f \) in row 6) due to the contraction of lower lying orbitals – this effect is called the lanthanide contraction.

Review: Why are the ‘s’, ‘p’, ‘d’ and ‘f’ element blocks so named?

Answer:
Wrap Up: Complete *Orbital ‘box’ diagrams, electronic configurations* and *‘dot’ diagrams* for the following:

<table>
<thead>
<tr>
<th>Atom</th>
<th>Electronic configuration</th>
<th>Orbital ‘box’ diagram</th>
<th>‘Dot’ diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>F</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>P</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ne</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Question: What similarities do you notice between the atoms’ respective ‘dot’ and ‘box’ diagrams?
Exceptions (take home assignment): Break down of the ‘dot’ models – ground and bound states

Task: Draw an orbital ‘box’ diagram and sketch a ‘dot’ diagram illustrating the electronic structure of carbon. Do you notice any inconsistencies?

C:  

\[ 1s \quad 2s \quad 2p \]

Carbon: the isolated atom’s ground state

Carbon: the bonded atom’s bound state

Answer:

The ‘dot’ and Lewis symbols represent the bonding configurations of the atoms’ electrons (i.e. what they are when the atom is bonded to other atom(s) - they have the max. number of unpaired e)

Orbital box, electronic configuration and energy level diagram representations show the lowest energy (or ground state) electron arrangement.

Elements in groups II, III and IV have different ground and bound electronic configurations

Task: Draw orbital ‘box’ diagrams illustrating the ground and bound states of Be

Be:  

\[ 1s \quad 2s \quad 2p \]

Ground state

Be:  

\[ 1s \quad 2s \quad 2p \]

Bound state

Question: Why do elements in groups II, III and IV have different ground and bound electronic configurations?

Answer: ‘you have to spend money to make money’ – what does this mean in terms of the ground and bound electronic states of atoms?
"Electron Addresses"

The following question was taken from your 3rd practice midterm:

Draw ground state orbital ‘box’ diagrams and write ground state electronic configurations for the following atoms and ions:

**Carbon atom**
Orbital ‘box’ diagram

Electronic Configuration

**Oxide anion**
Orbital ‘box’ diagram

Electronic Configuration

**Sodium atom**
Orbital ‘box’ diagram

Electronic Configuration

**Hydrogen atom**
Orbital ‘box’ diagram

Electronic Configuration

**Magnesium cation**
Orbital ‘box’ diagram

Electronic Configuration
Periodic Properties of the Elements

Background Discussion: What do we already know about the origins of today’s modern periodic table? What periodic trends do we already know?

Dmitri Mendeleev
“What was I doing back in 1869?”

Mendeleev’s Periodic Table

Mendeleev’s Predictions (all pretty good!)

Germanium (eka-silicon)

- Atomic mass: About 72 amu, 72.64 amu
- Density: 5.5 g/cm³, 5.35 g/cm³
- Formula of oxide: XO₂, GeO₂
- Formula of chloride: XCl₄, GeCl₄
Overview of Periodic Trends

Essentially all periodic trends follow the same general ‘bottom left to top right’ scheme. See generic diagram the below. The periodic trends examined will be:

- Electronegativity
- Ionization Energy
- Atomic Size (radius)*
- Electron Affinity

Recall: The layout of the Periodic Table is directly correlated to the elements’ electronic structures.

Additional chemical and physical trends among the table’s constituents can be understood by ‘popping the hood’ on these elements and determining the relationship between atomic properties electronic structures.

Question: Which ‘bottom left → top right trend’, mentioned above, have we already encountered? Is this a ‘real’ chemical trend?

Answer:
Trends in Electronegativity:

**Note:** Electronegativity is not a pure atomic property – it is a derived mathematically from *electron affinity* and *ionization energy* values.

**Discussion:** What atomic scale factors do you think effect how strongly the nucleus ‘pulls’ on its ‘orbiting’ electrons (so, in turn, effecting each atomic property)? **Hint:** consider the analogy of the Earth and its satellites.

1. 

2. 

3. 

230
**Effective Nuclear Charge ($Z_{\text{eff}}$): The real ‘Man behind the Curtain’**

**Basic definition of $Z_{\text{eff}}$:** The ‘pull’ orbiting *valence* electron(s) ‘feel’ from their respective positively charged nucleus, as modified by screening (*core*) electrons.

**Diagram**

Electrons outside have no effect on effective nuclear charge for electron of interest.

Electrons between electron of interest and nucleus cancels some of the positive nuclear charge.

**Equation:**

\[ Z_{\text{eff}} = Z - S \]

Effective nuclear charge

Actual nuclear charge

Charge screened by other electrons
Example: Lithium

$Lithium$: $1s^2 2s^1$

Fluorine: $1s^2 2s^2 2p^5$

**Facts:**

- As $Z$ (atomic number) increases ‘across a row’, the effect of shielding on the valence electrons AND their distance from the nucleus, remains $\sim$constant.

**Result:** Effective Nuclear Charge ($Z_{\text{eff}}$) increases for each atom across every row in the Periodic table.
Graph of $Z_{\text{eff}}$ vs Atomic Number

**Features of the Graph:**

- **Main**
- **Subtle**

**Atomic Radius**

**Discussion:** How and why do trends in $Z_{\text{eff}}$ effect the size (radius) of atoms ‘across a row’ and ‘down a column’ in the periodic table.

1. ‘Across a Row’

2. ‘Down a Column’
Atomic Radius Trends

The atomic radii of the main group elements follow the classic ‘bottom left → top right’ periodic trend.

The atomic radii of the main group elements follow the classic ‘bottom left → top right’ periodic trend.
Typical Question: Arrange the following atoms in order of increasing atomic radii: Na, Be, Mg.

<table>
<thead>
<tr>
<th></th>
<th>Be</th>
<th>Na</th>
<th>Mg</th>
</tr>
</thead>
</table>

Questions of this type (as well as for other periodic trends) often select three elements from the periodic table that have a ‘triangular’ relationship.

Understanding the classic ‘bottom left → top right’ periodic trend allows for the answer to be determined.

Answer:

Ionization Energy

Discussion: What is ionization? What then is 1st ionization energy?

1st Ionization Energy: Energy required to remove the first electron from a gaseous atom or ion.

Example: Sodium

\[ \text{Na} (g) \rightarrow \text{Na}^+ (g) + e^- ; \ I_1 = 496 \text{ kJ/mol} \]

Task: Draw electron dot diagrams illustrating this process
Discussion: How do you think trends in $Z_{eff}$ and atomic radius effect trends in 1$^{st}$ ionization energy ‘across a row’ and ‘down a column’ in the periodic table. *Who wins!*?

1. ‘Across a Row’

2. ‘Down a Column’

The 1$^{st}$ I.E. of the atoms generally follow the classic ‘bottom left → top right’ periodic trend*
Subtle Trends in 1st I.E.: How can subtle deviations (‘peaks’) from the general trend across any row be rationalized for the group II, III and group V, VI elements?

Task: Draw out ground state orbital ‘box’ diagrams for Be and B, as well as N and O. What differences do you notice between the two diagrams? How do these features correlate with the unexpected ‘peaks’.

‘Box’ Diagram for Be

Be:  
1s  2s  2p

‘Box’ Diagram for B

B:  
1s  2s  2p

‘Box’ Diagram for N

N:  
1s  2s  2p

‘Box’ Diagram for O

O:  
1s  2s  2p

Notes:

Removing electrons from full or half full (p or d) SUBSHELLS requires more energy than that required to remove electrons from ‘adjacent’ atom’s incomplete SUBSHELLS

Typical Question: Using only the periodic table as a guide, list the following atoms in order of increasing 1st ionization energy: O, N, Li, Na. Hint: consider both general and subtle trends.
Subsequent Ionization Energies

Discussion: What is the definition of second ionization energy ($I_2$)? Would you expect this value to be higher or lower than $I_1$ for Na? Why?

\[
\text{Na}^+ (g) \rightarrow \text{Na}^{2+} (g) + e^-; \quad I_2 = \text{_______ kJ/mol}
\]

Electron dot diagram:

| TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol) |
|-----------------------------------------------|-----------------------------------------------|
| **Element** | **IE$_1$** | **IE$_2$** | **IE$_3$** | **IE$_4$** | **IE$_5$** | **IE$_6$** | **IE$_7$** |
| Na          | 496        | 4560       |            |            |            |            |            |
| Mg          | 738        | 1450       | 7730       |            |            |            |            |
| Al          | 578        | 1820       | 2750       | 11,600     |            |            |            |
| Si          | 786        | 1580       | 3230       | 4360       | 16,100     |            |            |
| P           | 1012       | 1900       | 2910       | 4960       | 6270       | 22,200     |            |
| S           | 1000       | 2250       | 3360       | 4560       | 7010       | 8500       | 27,100     |
| Cl          | 1251       | 2300       | 3820       | 5160       | 6540       | 9460       | 11,000     |
| Ar          | 1521       | 2670       | 3930       | 5770       | 7240       | 8780       | 12,000     |

Notes:
Electron Affinity

Discussion: What is electron affinity?

**Electron Affinity**: Energy released when an electron is added to a gaseous atom or ion.

Example: Chlorine

\[
\text{Cl (g) + e}^- \rightarrow \text{Cl}^- (g) ; \Delta E = -349 \text{ kJ/mol}
\]

Task: Draw electron dot diagrams illustrating this process

Trends in Electron Affinity

*Electron affinity and electronegativity* are similar, except that electron affinity is measured experimentally, while electronegativity is determined mathematically.

Periodic style representations of atomic electron affinity trends. Note both the general and subtle features ‘across each row’
Discussion: Why do certain elements, such as Be, N, He and Ne, have negligible electron affinity values? **Hint:** draw out their respective ‘box’ diagrams

<table>
<thead>
<tr>
<th>Be: 1s 2s 2p</th>
<th>N: 1s 2s 2p</th>
</tr>
</thead>
<tbody>
<tr>
<td>He: 1s 2s 2p</td>
<td>Ne: 1s 2s 2p</td>
</tr>
</tbody>
</table>

**Answer:**

While *electron affinity* and *electronegativity* share a similar general periodic trend, ‘subtle’ subshell factors must also be taken into account with electron affinity.
"Trends"

The following question was taken from your 3rd practice midterm:

List the following properties of Li, K and Ne in order of:

Increasing atomic radius (smallest first)

Increasing effective nuclear charge, $Z_{\text{eff}}$ (smallest first)

Decreasing 1st ionization energy (largest first)
Appendix:

Trends in Metallic Character II

Trends in Atomic Radius
Radii of Atoms and Their Cations (pm)

<table>
<thead>
<tr>
<th>Group 1A</th>
<th>Group 2A</th>
<th>Group 3A</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>Li$^{+}$</td>
<td>Be</td>
</tr>
<tr>
<td>152</td>
<td>60</td>
<td>112</td>
</tr>
<tr>
<td>Na</td>
<td>Na$^{+}$</td>
<td>Mg</td>
</tr>
<tr>
<td>186</td>
<td>95</td>
<td>160</td>
</tr>
<tr>
<td>K</td>
<td>K$^{+}$</td>
<td>Ca</td>
</tr>
<tr>
<td>227</td>
<td>133</td>
<td>197</td>
</tr>
<tr>
<td>Rb</td>
<td>Rb$^{+}$</td>
<td>Sr</td>
</tr>
<tr>
<td>248</td>
<td>148</td>
<td>215</td>
</tr>
</tbody>
</table>

Radii of Atoms and Their Anions (pm)

<table>
<thead>
<tr>
<th>Group 6A</th>
<th>Group 7A</th>
</tr>
</thead>
<tbody>
<tr>
<td>O</td>
<td>O$^{2-}$</td>
</tr>
<tr>
<td>73</td>
<td>140</td>
</tr>
<tr>
<td>S</td>
<td>S$^{2-}$</td>
</tr>
<tr>
<td>103</td>
<td>184</td>
</tr>
<tr>
<td>Se</td>
<td>Se$^{2-}$</td>
</tr>
<tr>
<td>117</td>
<td>198</td>
</tr>
<tr>
<td>Te</td>
<td>Te$^{2-}$</td>
</tr>
<tr>
<td>143</td>
<td>221</td>
</tr>
</tbody>
</table>
Recap and overview: We have already investigated the structure of the atom in terms of ‘old style’ dot diagrams and Lewis symbols.

Recall the ‘battleship’ analogy with regard to # valence electrons an atom has (= element’s column in p. table) and the number of shells of electrons an atom has (= element’s row in p. table) an atom has.

Review question: Draw ‘dot’ diagrams and Lewis symbols for Li and F atoms

The Octet Rule (Full Valence Shell rule)

ATOMS WITH FULL OUTER (VALENCE) SHELLS ARE STABLE ⇒ Atoms will lose, gain or share electrons to have an inert gas (full valence shell) configuration.

THIS IS THE ‘DRIVING’ FORCE BEHIND ALL CHEMICAL PROCESSES.
1. Ionic bonding – the formation of LiF

Recall: Ionic bonds form between atoms (metal and non-metal, which then become ions) with a large difference in electronegativity.

Draw ‘dot’ diagrams illustrating the reaction between Li and F atoms to form LiF.

Dr. Mills’ favorite saying:

“Chemistry is a bit like Scottish soccer – it’s basically a bunch of round things bumping into one another”

2. Covalent bonding – the formation of F₂(g)

Recall: Covalent bonds form between atoms (two non-metals) with little or no difference in electronegativity.

Draw ‘dot’ diagrams illustrating the reaction between two F atoms to form F₂. Note: Lewis style diagrams of covalently bound molecules (e.g. F₂) are called Lewis structures.

One time British soccer legend ‘Gazza’ with a gyro

Gazza playing out his career with Glasgow Rangers
## Simple Lewis Structures

Overview: Lewis structures are electron ‘maps’ of molecules, which are in turn constructed from the Lewis symbols of the molecule’s component atoms.

The number of bonds an atom forms in a molecule (VALENCEY) = Number of UNPAIRED valence electrons it has as an atom

Task: Complete the following table

<table>
<thead>
<tr>
<th>Atom</th>
<th>Number valence electrons</th>
<th>Lewis Symbol</th>
<th>Valencey (number of bonds formed)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>N</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
**EZ Lewis Symbols** – think of an *unpaired* valence electron as ‘a hand that needs to be held’ (I could not think of a more masculine analogy!). Then just have the atoms ‘hold hands’ (form bonds by converting unshared e⁻ to shared pairs of e⁻) to make the required molecule’s Lewis structure. *H₂O example.*

**Task:** Write formal Lewis symbols and ‘EZ’ Lewis symbols for the following atoms:

<table>
<thead>
<tr>
<th>Atom</th>
<th>Formal Lewis Symbol*</th>
<th>‘EZ’ Lewis Symbol**</th>
<th>Valencey</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>N</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Remember: Just have the atoms, as represented by ‘EZ’ Lewis symbols, ‘hold hands’ to make the required molecule’s Lewis structure.
Task: Draw Lewis structures for the following molecules.

Remember: Each atom in a molecule must have as many bonds as its valence (number of unpaired electrons). **Double or Triple bonds often arise from applying this rule.**

Note: The total number of valence electrons in a Lewis structure is simply the sum of those ‘owned’ by each of the molecule’s component atoms. Write this information next to each of the above Lewis structures.

- **Ammonia (NH₃)**
  
  \[
  \begin{array}{c}
  \text{H} \\
  \text{N} \\
  \text{H}
  \end{array}
  \]

- **Water (H₂O)**
  
  \[
  \begin{array}{c}
  \text{H} \\
  \text{O} \\
  \text{H}
  \end{array}
  \]

- **Methane (CH₄)**
  
  \[
  \begin{array}{c}
  \text{H} \\
  \text{H} \\
  \text{H} \\
  \text{C}
  \end{array}
  \]

- **Phosphorus trichloride (PCl₃)**
  
  \[
  \begin{array}{c}
  \text{P} \\
  \text{Cl} \\
  \text{Cl} \\
  \text{Cl}
  \end{array}
  \]

- **Oxygen gas (O₂)**
  
  \[
  \begin{array}{c}
  \text{O} \\
  \text{O}
  \end{array}
  \]

- **Nitrogen Gas (N₂)**
  
  \[
  \begin{array}{c}
  \text{N} \\
  \text{N}
  \end{array}
  \]

- **Hydrogen Fluoride (HF)**
  
  \[
  \begin{array}{c}
  \text{H} \\
  \text{F}
  \end{array}
  \]

- **Dihydrogen monosulfide (H₂S)**
  
  \[
  \begin{array}{c}
  \text{H} \\
  \text{S} \\
  \text{H}
  \end{array}
  \]
Lewis Structures for More Complex Molecules - ‘The Rules’

Use the following rules to figure out the Lewis structure of ANY molecule (the above are simpler examples of the application of this ‘global’ set of rules)

Worked Example – Carbon Dioxide (CO$_2$)

1. Sum the valence electrons from all the atoms in the molecule or ion.

   For anions (-ve), ADD one e$^-$ per negative charge on the ion
   For cations (+ve), SUBTRACT one e$^-$ per positive charge on the ion

   ⇒ For CO$_2$:

2. Write the atoms on the page with the HIGHEST valencey atom in the center:

   ⇒ For CO$_2$:

3. Connect the outer atoms to the center atom with single lines(s) – these bonds (pair of shared electrons) are the minimum requirement for a molecule to exist.

4. Complete the valence shells of the ‘outside’ atoms to give them stable valence configurations.
5. Count up all the electrons in the structure and compare to the number required (from rule 1). Place any excess electrons on the center atom.

6. If the center atom does not have enough electrons for a complete valence shell, CONVERT ‘OUTSIDE’ LONE PAIR ELECTRONS TO DOUBLE BONDS. Remember that each outside atom’s valencey must also be obeyed.

7. Double-check the valencey of all atoms and the total number of electrons in the structure.

\[
\begin{array}{c}
\text{\textbullet\textbullet} \\
C \quad = \\
\text{\textbullet\textbullet}
\end{array}
\]

**Task**: Follow the above rules to construct Lewis structures for the following molecules and ions:

1. CH\textsubscript{2}O
2. CO\textsubscript{3}\textsuperscript{2-}

Is there more than one way to write the Lewis structure of the carbonate (CO\textsubscript{3}\textsuperscript{2-}) ion? Draw all possible versions.

These ‘different versions’ of the Lewis structure are called **resonance structures** (class demonstration). Resonance structures stabilize otherwise ‘impossible’ molecules and molecular ions.
Breaking the Octet (Full Valency Shell) Rule

The expectation that the valency of each atom in a Lewis structure is the same as its respective Lewis symbol is often violated for more complex molecules or molecular ions (e.g. \( \text{CO}_3^{2-} \)). This is because Lewis structures are empirical, i.e. they don’t completely model the real molecule in every detail.

Lewis structures can be ‘modified’ in order to account for these inconsistencies:

**Resonance forms**: Allow for the modeling of structures possessing atoms with apparent ‘lower’ or ‘higher’ valencies than expected.

**Octet expansion**: Allows for the modeling of center atoms with greatly expanded (numerically larger) valencies than expected. Such molecular models may also have a series of resonance forms.

Examples of structures modeled with **resonance forms**:

Follow the rules for drawing a regular Lewis structure - don’t worry if the atom(s) expected valencies are smaller or larger than expected - if two of more resonance structures can be drawn the model ‘works’. This is called resonance stability and leads to a stable, delocalized electronic structure in the ‘real’ molecule.

Examples:

1. \( \text{O}_3 \)
2. \( \text{NO}_2 \)
Environmental Concerns

LA smog (NO₂), mostly sourced from vehicle emissions. Note the brown color

The Ozone ‘hole’ over Antarctica

An Ozone action day sign warning of excess O₃

Discussion: How can there simultaneously be too much (ozone action days) and too little (Antarctica’s ozone ‘hole’) ozone in the atmosphere.
Examples of structures modeled with using **octet expansion**: 

Two criteria must be met for a center atom to *expand its octet* (increase its valencey):

1. The atom must be bonded to a highly electronegative atom, such as F, O or Cl (why)?

2. The atom must be in the 3rd row or lower of the periodic table (why?)

⇒ Atoms that undergo octet expansion are limited to P, S, Cl, Br and I (as well as several other lower p block non-metals)

Expansion can only occur within *orbitals possessing the same principle quantum number* (i.e. within the same ‘shell’). **Recall**: Have you seen something like this elsewhere already?

**Orbital box diagrams can be used to explain the increased valencey of the ‘lower p block’ elements.**

**Example**: Orbital ‘box’ diagrams for the valence shell of phosphorous

Unexpanded (‘regular’):  

<table>
<thead>
<tr>
<th>P: [Ne]</th>
<th>3s</th>
<th>3p</th>
<th>3d</th>
</tr>
</thead>
</table>

Expanded:  

<table>
<thead>
<tr>
<th>P: [Ne]</th>
<th>3s</th>
<th>3p</th>
<th>3d</th>
</tr>
</thead>
</table>

**Question**: What is the valencey of phosphorus with a ‘regular (unexpanded)’ and an expanded valence shell?
Task: Draw Lewis symbols for phosphorus with a ‘regular (unexpanded) and an expanded valence shell. Now draw ‘EZ’ Lewis structures for PCl\(_3\) (unexpanded Octet) and PCl\(_5\) (expanded octet)*

PCl\(_3\) \hspace{1cm} PCl\(_5\)

Electrons from the valence shell can be promoted to easily accessible (empty) 3d orbitals. This increases the number of unpaired valence electrons.

The total number of unpaired electrons in the expanded valence shell = the ‘new’ valencey of the atom. This can be expressed in the form of an ‘expanded’ Lewis symbol for the atom*.

*even though there are more than 8e in the expanded octet (after covalent bonds have been formed), the octet rule is still said to be obeyed as there are no remaining unpaired electrons in the valence shell of the center atom

Partial and complete octet expansion

Phosphorous (as shown above) can only promote one electron (from the 3s) to a 3d orbital in order to increase its valencey from 3 \(\rightarrow\) 5

Third (and later) row elements in groups VI – VIII have more than one ‘box’ of paired electrons in the ground state, so can increase their valencies even further by promoting more (initially paired) electron(s) from the \(s\) and \(p\) orbitals to the \(d\) orbitals

Electrons are sequentially promoted from the \(p\), then the \(s\), orbitals to the \(d\) orbitals to increase valencey. See below example.
Example: Write Lewis structures for \( \text{SO}_2 \) (partially expanded octet) and \( \text{SO}_4^{2-} \) (fully expanded octet).

**Trick**: Draw box diagrams first, and then construct an analogous ‘EZ’ or ‘rules’ Lewis structure.

**Ground State**:  
\[ \text{S: } [\text{Ne}] \quad \square \quad \square \quad \square \quad \square \quad \square \quad \square \quad 3s \quad 3p \quad 3d \]

**Partially expanded**:  
\[ \text{S: } [\text{Ne}] \quad \square \quad \square \quad \square \quad \square \quad 3s \quad 3p \quad 3d \]

**Fully expanded**:  
\[ \text{S: } [\text{Ne}] \quad \square \quad \square \quad \square \quad \square \quad 3s \quad 3p \quad 3d \]

**Additional Problems**: Draw Lewis structures for \( \text{XeF}_2, \text{XeF}_4 \) and \( \text{XeF}_6 \)
Formal Charge

Observation: We can now draw Lewis structures for molecular ions which, by definition (for anions), have ‘extra’ electrons. These ‘extra’ electrons give the ion its overall charge.

Fact: Where these ‘extra’ electrons are located in the Lewis structure is indicated by determining the *formal charge* of each atom in the structure.

The formal charge is the ‘charge’ on an atom in a Lewis structure, if it is assumed all the atoms in the structure have the same electronegativity (*yes, it’s another empirical model!*)

**Formal charges allow us to determine the ‘best’ Lewis structure**

As with drawing Lewis structures, follow a set series of empirical rules (shown below) to determine the final answer

Rules for assigning formal charge: Worked example - the cyanide ion

1. Construct a Lewis structure for the molecule or molecular ion.
2. determine the formal charge on each atom in the Lewis structure

The formal charge on an atom in a Lewis structure equals:

\[ \text{valence electrons it has as an isolated atom} \]

MINUS

\[ \text{# of valence electrons it ‘owns’ in the Lewis structure} \]

The sum of the formal charges for any molecule = zero

The sum of the formal charges for any molecular ion = overall ionic charge

For the cyanide ion:

Discussion: based on electronegativity trends, is the location of the electron in cyanide surprising? How does this compare with the hydroxide ion?

Example: Determine the formal charges on each atom in one resonance form of the CO$_3^{3-}$ ion’s Lewis structure.
The following question was taken from your 3rd practice midterm:

**Question 2a (20 points) Draw Lewis structure(s) for the PO$_4^{3-}$ ion, *include all possible resonance forms* and include formal charge labels on one of your structures. Assume a completely expanded octet for phosphorus.**
More Chemical Bonding

Reading: Ch 10: section 1 - 8  
Ch 9: section 4, 6, 10

Homework: Chapter 10: .31, 33, 35*, 39*, 43, 47, 49*  
Chapter 9: 43, 45, 55*, 57, 75*, 77, 79

* = ‘important’ homework question

Molecular Geometry – VSEPR Theory

The shape of most molecules in 3D can be determined by applying the Valence Shell Electron Pair Repulsion (VSEPR) Theory to a Lewis structure of the respective molecule.

Electron pairs in the valence shell of a center atom (as drawn in a Lewis structure) repel one another as they have the same negative charge.

The 3-D shape of a molecule is directly correlated to how the valence electron pairs are arranged in (3-D) in order to be as greatly separated from one another as possible.

<table>
<thead>
<tr>
<th>Number of Electron Domains</th>
<th>Arrangement of Electron Domains</th>
<th>Electron Domain Geometry</th>
<th>Predicted Bond Angles</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td></td>
<td>Linear</td>
<td>180°</td>
</tr>
<tr>
<td>3</td>
<td></td>
<td>Trigonal planar</td>
<td>120°</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td>Tetrahedral</td>
<td>109.5°</td>
</tr>
<tr>
<td>5</td>
<td></td>
<td>Trigonal-bipyramidal</td>
<td>120°, 90°</td>
</tr>
<tr>
<td>6</td>
<td></td>
<td>Octahedral</td>
<td>90°, 180°</td>
</tr>
</tbody>
</table>

Hybridization (later)  
Bond Types (later)
Consider each electron pair (bonded or lone) as ‘clumps’ of negative charge. These clumps adopt the above 3-D shapes in order to obey the VSEPR effect (see appendix and slide).

Examples of AXₙ structures: Draw Lewis structures and determine the 3-D molecular shapes of carbon dioxide (CO₂), methanal (CH₂O) and the sulfate anion (SO₄²⁻). Recall that you have previously completed Lewis structures for these species.
Molecular shape vs Electronic shape

The molecular (where the atoms are) and electronic (where the ‘clumps’ of electrons are) shapes of molecules are often different.

**Recall:** The valence electron pairs’ (bonded (X) and lone (E)) determine the overall electronic shape of the molecule.

**But:** The positions of the molecule’s atoms relative to one another (after the electronic shape has been fixed) determine the molecular shape.

**Examples:** Draw Lewis structures and determine the 3-D molecular and electronic shapes of methane (CH₄), water (H₂O) and ammonia (NH₃).
The relationship between molecular shape, electronic shape and numbers of bonding (X) and/or lone (E) pairs of valence electrons. See appendix.

### Table 10.1 Electron and Molecular Geometries

<table>
<thead>
<tr>
<th>Electron Groups*</th>
<th>Bonding Groups</th>
<th>Lone Pairs</th>
<th>Electron Geometry</th>
<th>Molecular Geometry</th>
<th>Approximate Bond Angles</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>Linear</td>
<td>Linear</td>
<td>$180^\circ$</td>
<td><img src="image1" alt="Example" /></td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>Trigonal planar</td>
<td>Trigonal planar</td>
<td>$120^\circ$</td>
<td><img src="image2" alt="Example" /></td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>1</td>
<td>Trigonal planar</td>
<td>Bent</td>
<td>$&lt;120^\circ$</td>
<td><img src="image3" alt="Example" /></td>
</tr>
<tr>
<td>4</td>
<td>4</td>
<td>0</td>
<td>Tetrahedral</td>
<td>Tetrahedral</td>
<td>$109.5^\circ$</td>
<td><img src="image4" alt="Example" /></td>
</tr>
<tr>
<td>4</td>
<td>3</td>
<td>1</td>
<td>Tetrahedral</td>
<td>Trigonal pyramidal</td>
<td>$&lt;109.5^\circ$</td>
<td><img src="image5" alt="Example" /></td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>2</td>
<td>Tetrahedral</td>
<td>Bent</td>
<td>$&lt;109.5^\circ$</td>
<td><img src="image6" alt="Example" /></td>
</tr>
</tbody>
</table>

Examples of 5 and 6 electron domain systems. See appendix and slides

#### 6 electron pairs

- $AX_6$
- $AX_5E$
- $AX_4E_2$

#### 5 electron pairs

- $AX_5$
- $AX_4E$
- $AX_3E_2$
- $AX_3E_3$

- Octahedral
- Square pyramidal
- Square planar
- Trigonal bipyramidal
- See-saw
- T-shaped
- Linear
Examples of AXmE_n structures: Draw Lewis structures and determine the 3D molecular and electronic shapes of water XeF_4 and NO_2. Recall that you have previously completed Lewis structures for these species.
Question of the week: Draw Lewis structures for the Chlorate ion (ClO$_4^-$). Determine their respective molecular and electronic shapes using VSEPR theory. Assume a completely expanded octet (valencey = 7) for chlorine.
Hybridization and Molecular Orbital (MO) Theory

Task: Read sections 10.6 - 10.8 of your text. Make brief notes on hybridization and MO theory. Reference the below diagrams in your notes

Hybridization Overview

The Organic Chemist’s Trick: An easy way to determine the hybridization of an atom is to simply count the number of electron domains (‘clumps’) it has - seen as either single, double, triple bonds and/or lone pairs. This relates directly to its respective orbital box diagram (hybridization).

<table>
<thead>
<tr>
<th>Hybridization (Carbon)</th>
<th>Number and Types of Bond (σ, π)</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>sp³</td>
<td>4 x σ bonds (or 3 x σ bonds + 1 l.p. for N) (4 single bonds)</td>
<td>![Example Image]</td>
</tr>
<tr>
<td>sp²</td>
<td>3 x σ bonds, 1 x π bond (2 single bonds, 1 double bond)</td>
<td>![Example Image]</td>
</tr>
<tr>
<td>sp</td>
<td>2 x σ bonds, 2 x π bond (2 double, or 1 triple and 1 single bonds)</td>
<td>![Example Image]</td>
</tr>
</tbody>
</table>

‘EZ’ Lewis

<table>
<thead>
<tr>
<th>Hybridization</th>
<th>#Domains</th>
<th>Shape</th>
<th>Orbital box Diagrams</th>
</tr>
</thead>
<tbody>
<tr>
<td>sp³</td>
<td>4</td>
<td>Td</td>
<td>X: 2s 2p sp³</td>
</tr>
<tr>
<td>sp²</td>
<td>3</td>
<td>Trig. Planar</td>
<td>X: 2s 2p sp² p²</td>
</tr>
</tbody>
</table>

*See appendix for sp² hybridization example
Case Study: $sp^3$ hybridization

Methane ($CH_4$) possesses 4 identical ‘single’ bonds.

Hybridization center atom: $sp^3$

Bond Type(s): $4 \times \sigma$
## Summary

<table>
<thead>
<tr>
<th>Number of Electron Groups</th>
<th>Electron Geometry (from VSEPR Theory)</th>
<th>Hybridization Scheme</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
<td>$sp$</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar</td>
<td>$sp^2$</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
<td>$sp^3$</td>
</tr>
<tr>
<td>5</td>
<td>Trigonal bipyramidal</td>
<td>$sp^3d$</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral</td>
<td>$sp^3d^2$</td>
</tr>
</tbody>
</table>

*Use the above information to complete the table on page 1 of this packet!*
The Strength of Covalent Bonds (Bond Enthalpy)

Background: Enthalpy the energy (in kJ) required to break one mole of a specified type of bond. Units are kJ/mole. For Cl₂:

\[
\begin{align*}
\text{Cl} = \text{Cl} & \rightarrow \text{Cl} + \text{Cl} ; \Delta H = 242 \text{ kJ/mol} \\
\end{align*}
\]

i.e. it takes 242 kJ of energy to break one mole of Cl-Cl bonds (see appendix)

Question: Which do you think are the stronger bonds? Why?

1. C-C (carbon–carbon single bond) or C=C (carbon–carbon double bond)

2. C-O (carbon–oxygen single bond) or C-C (carbon–carbon single bond)
Task: Work out the bond enthalpy for the C-H bond in methane (CH₄), given the following thermochemical equation:

\[ \text{CH}_4(\text{g}) \rightarrow \text{C(} \text{g}) \ + \ 4 \text{ H(g)} ; \Delta H = 1660 \text{ kJ/mol} \]

Hint: Draw a Lewis style bond breaking equation.

**Recall the Thermochemistry packet** – the DIFFERENCE in bond strengths (enthalpies) between reactants and products is ΔH for that process. If the bonds formed in the products are ‘stronger’ than those in the reactants the process is exothermic, i.e a –ve sign for ΔH.

**Remember:** A thermochemical equation assumes the same # moles of actual reactant and product are involved in the chemical process as indicated by their respective stoichiometric constants (balancing numbers)

Task: Use the bond enthalpy data in the appendix to determine ΔH for the combustion of methane
### Bond Enthalpy Trends

Describe the type of bonding in Cl₂, HCl and NaCl. Use this information to complete the following table:

<table>
<thead>
<tr>
<th>Compound</th>
<th>Cl-Cl</th>
<th>H-Cl</th>
<th>[Na]⁺[Cl]⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Model</strong></td>
<td><img src="image1" alt="Cl-Cl" /></td>
<td><img src="image2" alt="H-Cl" /></td>
<td><img src="image3" alt="Na⁺Cl⁻" /></td>
</tr>
<tr>
<td><strong>Bonding Type</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Bond Enthalpy</strong></td>
<td>243</td>
<td>431</td>
<td>787</td>
</tr>
</tbody>
</table>

**Discussion:** What trend do you notice with regard to the type of bonding and bond enthalpy?

### Bond Length / Strength Trends:

Complete the following table using the bond length and bond enthalpy tables (see appendix). **What correlation do you notice regarding the length and strength of chemical bonds?**

<table>
<thead>
<tr>
<th>Bond Coordinate</th>
<th>C-C</th>
<th>C-O</th>
<th>C=C</th>
<th>C=O</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Bond Length</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(pm)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Bond Enthalpy</strong></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(kJ/mole)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Notes:**
**Typical Question:** Arrange the following bond coordinates in order of increasing bond enthalpy and increasing bond length

1. C-N, C≡N, C=N
   
   Bond Enthalpy:
   
   Bond Length:

2. C-C, C-O, C-F
   
   Bond Enthalpy:
   
   Bond Length:

**Molecular Shape and Polarity**

**Discussion:** Would you expect the following molecules to have *net dipole moments*?

- **HCl**
  - Low electron density
  - High electron density
  - Polar bond

- **H₂O**
  - Net dipole moment

- **CO₂**
  - No net dipole moment
The Strength of Ionic Bonds (Lattice Energy)

**Definition:** The energy associated with forming a crystalline lattice from gaseous ions (see 9.4 and figure immediately below)

Since ionic bonds are electrostatic in nature and, by definition, omni-directional, *all* the respective +ve and –ve charges in the lattice will be attracted to one another. *Think of the lattice energy as the average of this force for a single pair of ions in the lattice*
Trends in Lattice Energy: Ion Size

**Discussion:** Using the accompanying figure as a guide, describe how relative size differences among ion pairs effect bond distance. How will this effect bond strength (lattice energy)?

*Check the lattice energy values on page 371 of Tro*

**Trends in Lattice Energy: Ion Charge**

**Discussion:** Describe how relative charge differences among the respective ion pairs, see figure below, effect bond strength (lattice energy). How would you expect the magnitude of this effect to compare with the above ion size trend?

*Check the lattice energy values on page 366 of Tro (charge beats size!!)*
The following question was taken from your 3rd practice midterm (you can now answer the second part also):

**Question 2a (20 points)** Draw Lewis structure(s) for the $PO_4^{3-}$ ion, include all possible resonance forms and include formal charge labels on one of your structures. *Assume a completely expanded octet for phosphorus.*

**Question 2b (5 points)** Use VSEPR theory to determine the electronic and molecular geometry of the phosphate ion:

**Electronic geometry:**

**Molecular geometry:**
The following question was taken from your 3rd practice midterm:

State whether the following bonds are considered to be covalent, polar covalent or ionic:

<table>
<thead>
<tr>
<th>Bond coordinate</th>
<th>Type of bond (covalent, polar covalent or ionic)</th>
</tr>
</thead>
<tbody>
<tr>
<td>O=O</td>
<td></td>
</tr>
<tr>
<td>O-H</td>
<td></td>
</tr>
<tr>
<td>Na-Cl</td>
<td></td>
</tr>
<tr>
<td>K-F</td>
<td></td>
</tr>
<tr>
<td>S-F</td>
<td></td>
</tr>
</tbody>
</table>
Appendix

(a) Linear geometry

(b) Trigonal planar geometry

Tetrahedral geometry

Tetrahedron

Trigonal bipyramidal geometry

Trigonal bipyramid

Octahedral geometry

Octahedron
The Continuum of Bond Types

Pure (nonpolar) covalent bond

Polar covalent bond

Ionic bond

Electrons shared equally

Electrons shared unequally

Electrons transferred

Electronegativity difference, ΔEN

Some Average Single- and Multiple-Bond Energies*

<table>
<thead>
<tr>
<th></th>
<th>Single Bonds</th>
<th>Multiple Bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>435</td>
<td>N==N 418</td>
</tr>
<tr>
<td>C</td>
<td>347</td>
<td>C==C 611</td>
</tr>
<tr>
<td>N</td>
<td>253</td>
<td>N==N 946</td>
</tr>
<tr>
<td>O</td>
<td>159</td>
<td>N==O 590</td>
</tr>
<tr>
<td>F</td>
<td>138</td>
<td>N==O (in O==O) 891</td>
</tr>
<tr>
<td>Si</td>
<td>184</td>
<td>C==O 837</td>
</tr>
<tr>
<td>P</td>
<td>190</td>
<td>C==O (as in H₂C==O) 745</td>
</tr>
<tr>
<td>S</td>
<td>215</td>
<td>O==O (in O₂) 498</td>
</tr>
<tr>
<td>Cl</td>
<td>213</td>
<td>C==O 1075</td>
</tr>
<tr>
<td>Br</td>
<td>251</td>
<td>Cl==C 121</td>
</tr>
<tr>
<td>I</td>
<td>251</td>
<td>Cl==N 127</td>
</tr>
</tbody>
</table>

Some Approximate Single and Multiple Bond Lengths*

<table>
<thead>
<tr>
<th></th>
<th>Single Bond Lengths</th>
<th>Multiple Bond Lengths</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>74</td>
<td>C==C 134</td>
</tr>
<tr>
<td>C</td>
<td>154</td>
<td>C==N 127</td>
</tr>
<tr>
<td>N</td>
<td>128</td>
<td>C==O 122</td>
</tr>
<tr>
<td>O</td>
<td>159</td>
<td>N==O 115</td>
</tr>
<tr>
<td>F</td>
<td>138</td>
<td>N==O (in O₂) 108</td>
</tr>
<tr>
<td>Si</td>
<td>184</td>
<td>Cl==C 121</td>
</tr>
<tr>
<td>P</td>
<td>213</td>
<td>Cl==N 127</td>
</tr>
<tr>
<td>S</td>
<td>213</td>
<td>Cl==O 115</td>
</tr>
<tr>
<td>Cl</td>
<td>251</td>
<td>Br==C 232</td>
</tr>
<tr>
<td>Br</td>
<td>251</td>
<td>Br==O 247</td>
</tr>
</tbody>
</table>

*In kilojoules per mole.
*In picometers (pm); 1 pm = 10⁻¹² m.
### TABLE 10.2 Common Cases of Adding Dipole Moments to Determine whether a Molecule Is Polar

<table>
<thead>
<tr>
<th>Nonpolar</th>
<th>Polar</th>
</tr>
</thead>
<tbody>
<tr>
<td>The dipole moments of two identical polar bonds pointing in opposite directions will cancel. The molecule is nonpolar.</td>
<td>The dipole moments of two polar bonds with an angle of less than 180° between them will not cancel. The resultant dipole moment vector is shown in red. The molecule is polar.</td>
</tr>
<tr>
<td><img src="image1.png" alt="Nonpolar Diagram" /></td>
<td><img src="image2.png" alt="Polar Diagram" /></td>
</tr>
<tr>
<td>The dipole moments of three identical polar bonds at 120° from each other will cancel. The molecule is nonpolar.</td>
<td>The dipole moments of four identical polar bonds in a tetrahedral arrangement (109.5° from each other) will cancel. The molecule is nonpolar.</td>
</tr>
<tr>
<td><img src="image3.png" alt="Nonpolar Diagram" /></td>
<td><img src="image4.png" alt="Nonpolar Diagram" /></td>
</tr>
<tr>
<td>The dipole moments of three polar bonds in a trigonal pyramidal arrangement (109.5° from each other) will not cancel. The resultant dipole moment vector is shown in red. The molecule is polar.</td>
<td></td>
</tr>
</tbody>
</table>

Note: In all cases where the dipoles of two or more polar bonds cancel, the bonds are assumed to be identical. If one or more of the bonds are different from the other(s), the dipoles will not cancel and the molecule will be polar.
\( \sigma \) and \( \pi \) bonds (sp\(^2\) hybridization) – CH\(_3\)O

3 x sp\(^2\) hybridized orbitals \( \rightarrow \) 3 x \( \sigma \) bonds

1 x p orbital \( \rightarrow \) 1 x \( \pi \) bond

Lewis structure  Valence bond model
Intermolecular Forces

**Background**

**Discussion:** What is the difference between an *intermolecular* force and an *intra*molecular force? **Hint:** Think about the difference between flying to Cleveland and Flying to Europe

*Intramolecular Force:*

*Intermolecular Force:*

**Example:** water
Types of Intermolecular Forces (weak bonds between molecules)

Intermolecular forces are what hold molecular materials together in the liquid or solid state (gases experience no intermolecular forces so are free to fill the container in which they are placed).

Intermolecular bonds are broken when energy (heat) greater than the intermolecular bond strength is applied to the material. This is why materials have specific melting and freezing points.

States of Matter and Heating / Cooling Curves

Notes:
Overview: There are THREE types of intermolecular force (bond):

<table>
<thead>
<tr>
<th>Type of Force</th>
<th>Strength</th>
<th>Notes</th>
</tr>
</thead>
<tbody>
<tr>
<td>London Dispersion Forces</td>
<td>weak - strong</td>
<td>Common to all molecular</td>
</tr>
<tr>
<td>(induced dipole - dipole)</td>
<td></td>
<td>materials</td>
</tr>
<tr>
<td>Dipole - Dipole</td>
<td>strong</td>
<td>Only for polar molecules</td>
</tr>
<tr>
<td>Hydrogen Bonding</td>
<td>very strong</td>
<td>Only for specific molecules</td>
</tr>
</tbody>
</table>

London Dispersion Forces (induced dipole – dipole bonding)

Theory:

1. Short lived *time dependant dipoles* are being created in atoms (and molecules) continually as electrons move around their respective orbital(s). Recall that a dipole is a special separation of charge. Since all atoms and molecules contain electrons, they all do this.

2. At close to the condensation point (gas - liquid), the atoms or molecules are moving slowly enough for an *induced dipole* to form between adjacent atoms or molecules. This spreads throughout the material, turning it to a liquid (or solid).
Analogy: Induced Dipole interactions are much like the ‘wave’ - seen at various sporting events when the crowd becomes ‘bored’ (like at Sox games).

CLASS DEMO: ‘Helium in the house’

Likely Exception: British soccer – extreme boredom

“Com’on lads, let’s see how they like the taste of this pointy metal fence”

The strength of London Dispersion Forces

Discussion: What basic property of an atom or molecule results in the formation of induced dipole – dipole bonding (London forces)? How then can the degree dipole – dipole bonding be increased? What macroscopic affect would this have?
The strength of an induced dipole – dipole bond is proportional to the number of electrons an atom or molecule has. Since atomic mass scales with the number of electrons:

**Strength of London of Forces ∝ Molecular mass ∝ boiling point***

*for atoms and molecules that only have induced dipole-dipole intermolecular forces

Boiling Points, # electrons and Molar masses ($M$) for the Nobel gases

<table>
<thead>
<tr>
<th>Nobel Gas</th>
<th># electrons</th>
<th>$M$ (g/mole)</th>
<th>Bpt. (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>2</td>
<td>4.0</td>
<td>4.6</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td>20.2</td>
<td>27.3</td>
</tr>
<tr>
<td>Ar</td>
<td>18</td>
<td>39.9</td>
<td>87.5</td>
</tr>
<tr>
<td>Kr</td>
<td>36</td>
<td>83.8</td>
<td>120.9</td>
</tr>
</tbody>
</table>

Molecular shape considerations

Discussion: Pentane (a) and isopentane (b) have identical molecular weights and molecular formulas. However, their shapes and boiling points are different. Explain.

(a) Pentane (bp = 309.4 K)  
(b) 2,2-Dimethylpropane (bp = 282.7 K)
Dipole-Dipole and Dipole – Ion interaction

A number of molecules have permanent dipoles, so experience stronger dipole–dipole interactions in addition to London dispersion forces.

Recall: Polar molecules have a net dipole (separation of charge). HCl is a good example of a polar molecule.

The \( \delta^+ \) and \( \delta^- \) ‘ends’ of polar molecules are attracted to one another – this is a dipole- dipole intermolecular force.

Example: HCl
Any molecule with a *permanent dipole* will undergo *dipole-dipole intermolecular bonding*. **Example**, \(\text{CH}_3\text{COCH}_3\) (polar C=O bond)

The strength of a dipole-dipole intermolecular interaction is related to the strength of a molecule’s permanent dipole (dipole moment).

**Strength of dipole-dipole force \(\propto\) Dipole moment \(\propto\) boiling point**

Boiling Points, Dipole moments for some similar Mwt. \(M\) molecules

(see appendix / slide)

<table>
<thead>
<tr>
<th>Compound</th>
<th>Formula</th>
<th>(M) (g/mole)</th>
<th>Dipole moment (\mu)</th>
<th>Bpt. (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Propane</td>
<td>(\text{CH}_3\text{CH}_2\text{CH}_3)</td>
<td>44</td>
<td>0.1</td>
<td>231</td>
</tr>
<tr>
<td>Dimethyl ether</td>
<td>(\text{CH}_3\text{OCH}_3)</td>
<td>46</td>
<td>1.3</td>
<td>254</td>
</tr>
<tr>
<td>Ethylene oxide</td>
<td>((\text{CH}_2)_2\text{O})</td>
<td>44</td>
<td>1.9</td>
<td>284</td>
</tr>
<tr>
<td>Acetaldehyde</td>
<td>(\text{CH}_3\text{CHO})</td>
<td>44</td>
<td>2.7</td>
<td>294</td>
</tr>
<tr>
<td>Acetonitrile</td>
<td>(\text{CH}_3\text{CN})</td>
<td>41</td>
<td>3.9</td>
<td>355</td>
</tr>
</tbody>
</table>
Hydrogen Bonding

Hydrogen bonding is a ‘special’ form of strong dipole-dipole interaction.

Hydrogen bonds are the strongest form of intermolecular force. A hydrogen bond is ~10% the strength as an intramolecular covalent bond.

Requirements of a hydrogen bond: the -X:δ⁻ • • • δ⁺H − δ⁻Y- coordinate

Diagram

Typical H-bond coordinates
Examples:

- Water (H₂O)
- Ammonia (NH₃)
- Ethanol
- DNA Base Pairs (see appendix)

Thymine

Adenine

Cytosine

Guanine
Hydrogen bonding greatly increases the boiling points of H-bonded materials. See figure.

Discussion: ‘Amazing water’: Based on the above graph, what’s the projected boiling point of water based only on induced dipole-dipole forces? Why is the actual boiling point of water so high? What consequences does this have for, oh, let’s say, the emergence of life on earth??
Discussion: Table salt (NaCl) is very soluble in water, while ‘oil’ (e.g. pentane) and water are immiscible – what types of intermolecular interaction(s) are responsible for these facts? (See appendix).

Recall the Chemist’s stock phrase regarding solubility / immiscibility:

“LIKE DISSOLVES IN LIKE”

Dipole – Ion interactions: NaCl (aq)

Polar- Polar vs non-Polar – non Polar interactions: ‘oil and water’
Summary

- All materials have induced dipole – dipole / London Dispersion forces (they all have electrons)
- Additional permanent dipole – dipole or H- bonding interactions occur for a small subset of molecules with the necessary molecular features
- H-bonded materials have much greater boiling points that predicted using only London dispersion force trends (see above figure)

**H-bonding >> Dipole - Dipole > London dispersion**

*strongest → weakest*

- Recall that ‘like dissolves in like’ due to complementary intermolecular forces (and vice versa):

These forces may contribute to or oppose the formation of a solution.
Appendix

Boiling Point vs Dipole moment for similar Mwt. Compounds

- Propane: CH₃CH₂CH₃, 44.09 g/mol
- Dimethyl ether: CH₃OCH₃, 46.07 g/mol
- Ethylene oxide: (CH₂)₂O, 44.05 g/mol
- Acetaldehyde: CH₃CHO, 44.05 g/mol
- Acetonitrile: CH₃CN, 41.05 g/mol

Boiling point (K) vs Dipole moment (D)
DNA replication
Exam Tips and Final Review

Ions

We committed a number of ion names and formulas to memory – the anions possess either –ide (mostly atomic anions, such as chloride, Cl\(^-\)) or –ate (molecular anions, such as SO\(_4^{2-}\), sulfate) suffixes. There are also many ‘in between’ molecular anions containing fewer, or occasionally more, oxygen atoms than the -ate ions. For example (from p 142):

<table>
<thead>
<tr>
<th>Ion formula</th>
<th>Name or ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl(^-)</td>
<td>chloride</td>
</tr>
<tr>
<td>ClO(^-)</td>
<td>Hypochlorite</td>
</tr>
<tr>
<td>ClO(_2^-)</td>
<td>Chlorite</td>
</tr>
<tr>
<td>ClO(_3^-)</td>
<td>Chlorate</td>
</tr>
<tr>
<td>ClO(_4^-)</td>
<td>Perchlorate</td>
</tr>
</tbody>
</table>

The above table is worth memorizing, as there is likely to be one or two questions on the final regarding the ‘in between’ molecular anions of oxygen and chlorine.

Example: Lithium chlorite has the formula:

a. LiClO  
b. LiClO\(_3\)  
c. LiClO\(_2\)  
d. LiClO\(_4\)  
e. LiCl

Solubility

A solubility chart will NOT be provided for final, although you will have access to a periodic table. This is not a problem, however, as there are only two basic facts to remember:

Chlorides are mostly soluble - AgCl(s) is an important exception

Sulfates are mostly soluble – BaSO\(_4\) (s) is an important exception

Use the above information to answer solubility questions on the final
Final Exam Review

Information

Your ACS standardized final exam is a comprehensive, 70 question multiple choice (a – d) test. Questions are graded as either correct or incorrect. No points are subtracted for wrong guesses. There are two versions of the test, so your neighbors will have a different version of the test.

I normalize your final exam score out of 70 to a score out of 200. This score out of 200 is included in your final course total.

Tips

A great way to ensure a good grade on your 101 and/or 102 final is to practice ACS style questions*. To help you with this, a pair of ACS study guides has been put on reserve in the library. Additionally, if you are going to take the MCAT or PCAT, the ACS guide is makes for an excellent resource for these tests' respective chemistry sections. If you wish to purchase a copy ($12 ea.) log on to:
http://www3.uwm.edu/dept/chemexams/guides/details_guides.cfm?ID=162

*Ask me to tell you the ‘BB’ story

Due to the number of questions set and the time allowed, most of the multiple choice questions you will meet on the final may be considered to be ‘lite’ versions of my midterm and quiz questions. The following tips will help you record a better score on your final:

1. The test is cumulative, so review everything we have covered since the beginning for the course.

2. Review all the topics, but concentrate on topics you have had difficulty with. Since the questions are not ‘super hard’, this will increase your number of correct answers. Do not fall in to the trap of studying what you are good at (you’ll get those questions right regardless, most likely), so preferentially study what you are ‘bad’ at.
3. Try to answer the questions in order when using a scantron sheet. It is better to guess a wrong answer (and then come back to it) than risk systematically filling out ovals ‘a line out’.

4. Work out the answers on the scratch paper provided, then check the possible answers provided. This will cut down on ‘red herring’ type errors (see below).

5. Watch out for obvious ‘red herrings’, as illustrated by the following example. MOST questions DO NOT have a red herring, but a reasonable fraction do:

Example: CO is the formula for:

a. copper  c. cobalt
b. carbon monoxide  d. Monocarbon monoxide

Sample Final Exam Questions

1. In all neutral atoms, there are equal numbers of:

a. electrons and protons  c. electrons and neutrons
b. protons and neutrons  d. electrons and positrons

2. Which pair of particles has the same number of electrons:

a. F, Mg$^{2+}$  c. P$^{3-}$, Al$^{3+}$
b. Ne, Ar  d. Br$^-$, Se

3. What is the mass percent of oxygen in Ca(NO$_3$)$_2$?

a. 29.3 %  c. 58.5 %
b. 47.1 %  d. 94.1%
4. A 24.0 g sample of carbon contains how many atoms:

(a) $6.02 \times 10^{23}$
(b) $1.20 \times 10^{24}$
(c) $3.01 \times 10^{23}$
(d) $2.04 \times 10^{24}$

5. When 1.187 g of a metallic oxide is reduced with excess hydrogen, 1.054 g of the metal is produced. What is the metallic oxide?

(a) $\text{Ag}_2\text{O}$
(b) $\text{Cu}_2\text{O}$
(c) $\text{K}_2\text{O}$
(d) $\text{Tl}_2\text{O}$

6. A single molecule of a certain compound has a mass of $3.4 \times 10^{-22}$ g. Which value comes closest to the mass of a mole of this compound?

(a) 50 g
(b) 100 g
(c) 150 g
(d) 200 g

7. The electronic configuration for the Ca atom is:

(a) $1s^22s^22p^63s^23p^64s^23d^2$
(b) $1s^22s^22p^63s^2$
(c) $1s^22s^22p^63s^23p^64s^2$
(d) $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$

8. Two moles of any gas will occupy what volume at STP?

(a) 22.4 L
(b) 11.2 L
(c) 4.48 L
(d) 44.8 L

Answers:

1. a. 5. b.
2. a. 6. d.
3. c. 7. c.
4. b. 8. d.
Post-Final Wrap Up

General chemistry final exams are graded immediately after they have been completed by the students. The final exam scores (out of 200), as well as overall course scores and letter grades, will be available from 10:00 am on Thursday of exam week. **Students can check their scores by sending Dr. Mills an e-mail request at any time before noon on Thursday of exam week.** In order to ensure confidentiality, students requesting such feedback must include the following code word(s) within their e-mail requests:

So Close, Yet so Far?

Unfortunately, it is sometimes the case that students find themselves just a few points below the C/D (55%) cut-off line after the completion of all course materials. In order for such students to achieve a passing ‘C’ grades, an optional 25 pt. extra credit assignment may be completed. **Students may only complete this assignment if they are in good academic standing (no more than one missed quiz, exam or lab) and contact Dr. Mills, via e-mail with a grade request, no later than noon on Thursday of exam week.** Such students’ final scores must have fallen no more than 25 points below the C/D cut-off in order for them to be eligible to take the assignment.

Dr. Mills will supply qualifying students with a copy of the extra credit assignment, as an e-mail attachment, via return e-mail. **The hard deadline for completing this assignment is 10:00 am on Friday of exam week – no exceptions.**
Chemistry 101: 1st Midterm Examination

Practice Questions
(see the class notes for answers to these questions)

Answer all four questions. Each question is worth 25 points. Please ensure you have all five pages of questions, as well as a formula sheet and periodic table before starting work. Only attempt the extra credit question after you have completed the four assigned problems. For numerical answers, include the correct number of significant figures and appropriate S.I. unit(s). For full credit you must….

Show all work

<table>
<thead>
<tr>
<th>Question</th>
<th>Score</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
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</tr>
<tr>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
</tr>
<tr>
<td>Extra Credit</td>
<td></td>
</tr>
</tbody>
</table>

Total

301
“The Wire”

Question 1a: A spool of copper (Cu) wire has a mass of 2.00 pounds and a diameter of 100 μm. Determine the wire’s mass, volume and length in the units specified below. Include any appropriate decimal prefixes in your final answers. Assume density copper (Cu) = 8.95 g cm$^{-3}$

Mass of the wire in kg:

Volume of the wire in cm$^3$:

Length of the wire in meters:

Question 1b: Write the complete atomic symbol for the isotope that contains 29 protons and 34 neutrons.

Complete atomic symbol:
“Ionic”

Question 2: Write the formulas and names of nine ionic compounds that may be formed through combining the anions and cations ions listed immediately below.

<table>
<thead>
<tr>
<th>Ionic Formula</th>
<th>Name of Ionic Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
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<tr>
<td></td>
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<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>
"Balance"

**Question 3:** Balance the following chemical equations:

a. The burning of liquid butane \((\text{C}_4\text{H}_{10}(l))\) in air

b. The neutralization of battery acid (sulfuric acid solution) with caustic soda (sodium hydroxide solution)

c. The reaction of solid diphosphorus pentoxide with water to form aqueous phosphoric acid

d. The decomposition of chalk \((\text{CaCO}_3)\), when heated, to form solid calcium oxide and carbon dioxide gas

e. The reaction of metallic zinc with aqueous sulfuric acid to form aqueous zinc (II) sulfate and hydrogen gas

**Extra credit:** State which general class of reaction each of the above belongs to
“% Mass”

Question 4: Calculate the % mass of each type of atom in the following materials:

BaSO₄

NO
Extra Credit

Expect a descriptive style question taken from the reading.
Data sheet

Density = mass/volume  1 kg = 2.205 lb  1 cm³ = 1 mL = 1 x10⁻⁶ m³
Density Lead (Pb) = 11.34 g cm⁻³  1 inch = 2.54 cm  1 mile = 1.6039 km
1 a.m.u. = 1.6606 x 10⁻²⁴ g  1 ft = 12 inches (exactly)  1 gallon = 3.786 L
Volume cylinder = πr²h  Volume sphere = (4πr³)/3
Diameter = 2r  1 pound = 16 ounces  Nₐ = 6.02 x10²³

Common Decimal Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Exponential Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Giga</td>
<td>G</td>
<td>10⁹</td>
</tr>
<tr>
<td>Mega</td>
<td>M</td>
<td>10⁶</td>
</tr>
<tr>
<td>Kilo</td>
<td>k</td>
<td>10³</td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>10⁻¹</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>10⁻²</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>10⁻³</td>
</tr>
<tr>
<td>micro</td>
<td>µ</td>
<td>10⁻⁶</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>10⁻⁹</td>
</tr>
</tbody>
</table>

Solubility Rules

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing NO₃⁻</td>
<td>None</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
</tr>
<tr>
<td>Br⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
</tr>
<tr>
<td>I⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>Ba²⁺, Hg²⁺, Pb²⁺</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Insoluble Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing CO₃⁻</td>
<td>NH₄⁺ &amp; group IA cations</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>NH₄⁺ &amp; group IA cations</td>
</tr>
<tr>
<td>OH</td>
<td>group IA cations</td>
</tr>
<tr>
<td>Ca²⁺, Sr²⁺, Ba²⁺ &amp; NH₄⁺</td>
<td></td>
</tr>
</tbody>
</table>

307
Chemistry 101: 2nd Midterm Examination

Practice Questions
(see the class notes for answers to these questions)

Answer all four questions. Each question is worth 25 points. Please ensure you have all five pages of questions, as well as a formula sheet and periodic table before starting work. Only attempt the extra credit question after you have completed the four assigned problems. For numerical answers, include the correct number of significant figures and appropriate S.I. unit(s). For full credit you must….

Show all work

<table>
<thead>
<tr>
<th>Question</th>
<th>Score</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
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</tr>
<tr>
<td>4</td>
<td></td>
</tr>
<tr>
<td>Extra Credit</td>
<td></td>
</tr>
<tr>
<td>Total</td>
<td></td>
</tr>
</tbody>
</table>
“Take out?”

**Question 1**: Monosodium glutamate (MSG) is a popular flavor enhancer used in many fast foods. Assuming MSG contains 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.6% Na, and has a molar mass of 169 g/mol determine:

The **Empirical formula of MSG**

The **Molecular formula of MSG**
"Equations and Solubility"

Question 2: Write balanced, complete and net ionic equations illustrating the reaction between aqueous solutions of silver nitrate and sodium phosphate. *Include all state symbols.*

Balanced chemical equation:

Complete ionic equation:

Net ionic equation:

List the names and formulas of five insoluble ionic compounds containing the hydroxide ion.
“Limiting”

Question 3: Lithium and nitrogen react to produce lithium nitride:

\[ 6 \text{Li(s)} + \text{N}_2 (g) \rightarrow 2 \text{Li}_3\text{N (s)} \]

If 7.00 grams of each reactant undergo a reaction with 85.0 % yield, how many grams of \( \text{Li}_3\text{N} \) are obtained from the reaction?
"Lewis symbols and Dot structures"

**Question 4:** Draw Lewis symbols and complete ‘dot’ structures and for the following:

<table>
<thead>
<tr>
<th>Lewis symbol</th>
<th>Dot structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon atom</td>
<td></td>
</tr>
<tr>
<td>Oxide anion</td>
<td></td>
</tr>
<tr>
<td>Sodium atom</td>
<td></td>
</tr>
<tr>
<td>Hydrogen atom</td>
<td></td>
</tr>
<tr>
<td>Magnesium ion</td>
<td></td>
</tr>
</tbody>
</table>
Extra Credit

Expect a descriptive style question taken from the reading.
**Data sheet**

Density = mass/volume
Density copper (Cu) = 8.95 g cm\(^{-3}\)
1 a.m.u. = 1.6606 x 10\(^{-24}\) g
Volume cylinder = \(\pi r^2 h\)

1 kg = 2.205 lb
1 inch = 2.54 cm
1 mile = 1.6039 km
1 ft = 12 inches (exactly)
1 gallon = 3.786 L
1 cm\(^3\) = 1 mL = 1 x 10\(^{-6}\) m\(^3\)

**Common Decimal Prefixes**

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</tr>
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<td>deci</td>
<td>d</td>
<td>10(^{-1})</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>10(^{-2})</td>
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<tr>
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<td>m</td>
<td>10(^{-3})</td>
</tr>
<tr>
<td>micro</td>
<td>(\mu)</td>
<td>10(^{-6})</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>10(^{-9})</td>
</tr>
</tbody>
</table>

**Solubility rules:**

<table>
<thead>
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<th>Soluble Compounds</th>
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</tr>
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<tbody>
<tr>
<td>Compounds containing NO(_3^{-})</td>
<td>None</td>
</tr>
<tr>
<td>Cl(^{-})</td>
<td>Ag(^{+}), Hg(^{2+}), Pb(^{2+})</td>
</tr>
<tr>
<td>Br(^{-})</td>
<td>Ag(^{+}), Hg(^{2+}), Pb(^{2+})</td>
</tr>
<tr>
<td>I(^{-})</td>
<td>Ag(^{+}), Hg(^{2+}), Pb(^{2+})</td>
</tr>
<tr>
<td>SO(_4^{2-})</td>
<td>Ba(^{2+}), Hg(^{2+}), Pb(^{2+})</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Insoluble Compounds</th>
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<td>Compounds containing CO(_3^{3-})</td>
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<td>PO(_4^{3-})</td>
<td>NH(_4^{+}) &amp; group IA cations</td>
</tr>
<tr>
<td>OH</td>
<td>group IA cations</td>
</tr>
<tr>
<td>Ca(^{2+}), Sr(^{2+}), Ba(^{2+}) &amp; NH(_4^{+})</td>
<td></td>
</tr>
</tbody>
</table>

314
Chemistry 101: 3rd Midterm Examination

Practice Questions
(see the class notes for answers to these questions)

Answer all four questions. Each question is worth 25 points. Please ensure you have all four pages of questions, as well as a formula sheet and a copy of the periodic table before starting work. For numerical answers, include the correct number of significant figures and appropriate S.I. unit(s). For full credit you must….

Show all work

<table>
<thead>
<tr>
<th>Question</th>
<th>Score</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
</tr>
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<tr>
<td>4</td>
<td></td>
</tr>
<tr>
<td>Total</td>
<td></td>
</tr>
</tbody>
</table>
"Electron Addresses"

Question 1 (25 points) Draw ground state orbital 'box' diagrams and write ground state electronic configurations for the following atoms and ions:

**Carbon atom**

Orbital ‘box’ diagram

Electronic Configuration

**Oxide anion**

Orbital ‘box’ diagram

Electronic Configuration

**Sodium atom**

Orbital ‘box’ diagram

Electronic Configuration

**Hydrogen atom**

Orbital ‘box’ diagram

Electronic Configuration

**Magnesium cation**

Orbital ‘box’ diagram

Electronic Configuration
“Lewis”

**Question 2a (20 points)** Draw Lewis structure(s) for the PO$_4^{3-}$ ion, include all possible resonance forms and include formal charge labels on one of your structures. Assume a completely expanded octet for phosphorus.

**Question 2b (5 points)** Use VSEPR theory to determine the electronic and molecular geometry of the phosphate ion:

**Electronic geometry:**

**Molecular geometry:**
"Trends"

Question 3a (15 points) List the following properties of Li, K and Ne in order of:

Increasing atomic radius (smallest first)

Increasing effective nuclear charge, \( Z_{\text{eff}} \) (smallest first)

Decreasing 1\textsuperscript{st} ionization energy (largest first)

Question 3b (10 points) State whether the following bonds are considered to be covalent, polar covalent or ionic:

<table>
<thead>
<tr>
<th>Bond coordinate</th>
<th>Type of bond (covalent, polar covalent or ionic)</th>
</tr>
</thead>
<tbody>
<tr>
<td>O=O</td>
<td></td>
</tr>
<tr>
<td>O-H</td>
<td></td>
</tr>
<tr>
<td>Na-Cl</td>
<td></td>
</tr>
<tr>
<td>K-F</td>
<td></td>
</tr>
<tr>
<td>S-F</td>
<td></td>
</tr>
</tbody>
</table>
**“Raft”**

**Question 4 (25 points)** The following reaction between calcium hydride and water is used to inflate life rafts and weather balloons:

\[
\text{CaH}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca(OH)}_2(\text{aq}) + 2 \text{H}_2(\text{g})
\]

If 47.0 grams of CaH_2(s) is completely reacted with an excess of water, then:

C. **How many moles of hydrogen gas will be produced?**

D. **What volume (in L) would the hydrogen gas generated in part (a) occupy at 15.0°C and 725 Torr?**
Extra Credit

Expect a descriptive style question taken from the reading.
Data sheet

Density = mass/volume

Density copper (Cu) = 8.95 gcm⁻³

1 a.m.u. = 1.6606 x 10⁻²⁴ g

Volume cylinder = πr²h

1 atm = 760 torr = 101 kPa

1 kg = 2.205 lb

1 inch = 2.54 cm

1 ft = 12 inches (exactly)

1 mile = 1.6039 km

1 gallon = 3.786 L

R = 0.0821 Latm/molK

Common Decimal Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Exponential Notation</th>
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</thead>
<tbody>
<tr>
<td>Giga</td>
<td>G</td>
<td>10⁹</td>
</tr>
<tr>
<td>Mega</td>
<td>M</td>
<td>10⁶</td>
</tr>
<tr>
<td>Kilo</td>
<td>k</td>
<td>10³</td>
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<tr>
<td>deci</td>
<td>d</td>
<td>10⁻¹</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>10⁻²</td>
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<tr>
<td>milli</td>
<td>m</td>
<td>10⁻³</td>
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<tr>
<td>micro</td>
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<td>10⁻⁶</td>
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<tr>
<td>nano</td>
<td>n</td>
<td>10⁻⁹</td>
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Solubility rules:

<table>
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<tr>
<th>Compounds containing</th>
<th>NO₃⁻</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td>None</td>
</tr>
<tr>
<td>Br⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td></td>
</tr>
<tr>
<td>I⁻</td>
<td>Ag⁺, Hg²⁺, Pb²⁺</td>
<td></td>
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<tr>
<td>SO₄²⁻</td>
<td>Ba²⁺, Hg²⁺, Pb²⁺</td>
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<table>
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<tr>
<th>Compounds containing</th>
<th>CO₃²⁻</th>
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<tbody>
<tr>
<td>NH₄⁺ &amp; group IA cations</td>
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<tr>
<th>Compounds containing</th>
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<tr>
<td>NH₄⁺ &amp; group IA cations</td>
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<thead>
<tr>
<th>Compounds containing</th>
<th>OH⁻</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>group IA cations Ca²⁺, Sr²⁺, Ba²⁺</td>
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<th>Exceptions</th>
</tr>
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<tr>
<td>NH₄⁺ &amp; group IA cations</td>
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