How to prepare for exams

1. Understand EXAMPLES in chapter(s)
2. Work PRACTICE EXERCISES
3. Work one problem from each class of problems at end of chapter
4. Answer as many questions as time permits from text web:  www.prenhall.com/brown
5. Test yourself with old exam from my web page:  SEE SYLLABUS

Review

Chapters 10, 11, 13

Chapter 10 - GASES

• Pressure
• What can we measure?
• The Gas Laws
• The Ideal Gas Equation
• Partial Pressures, Molecular Weight & Density
• Kinetic Molecular Theory
• Real Gases

PRESSURE IS THE FORCE ACTING ON AN OBJECT PER UNIT AREA

\[ P = \frac{F}{A} \]

MEASUREMENT OF PRESSURE

• The relationship between pressure (P) and the height (h) of a liquid in a barometer is

\[ P = g d h \]

where \( g \) => constant acceleration of gravity and \( d \) => density of liquid in manometer
**WHAT CAN WE MEASURE ?**

- **PRESSURE** of a gas
- **TEMPERATURE** of a gas
- **VOLUME** of a gas
- **AMOUNT** of a gas

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**Standard (STP) Conditions**

- **Standard Temperature** = 0°C = 273.15 K
- **Standard Pressure** = 1 atm

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**OTHER USEFUL VALUES at STP**

- “standard” volume = 22.4 Liters
- When n = 1 mole

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**Review of Gas Laws**

- **Boyle’s Law** \( P V = a \text{ constant} \)
- **Charles’s Law** \( \frac{V}{T} = a \text{ constant} \)
- **Avogadro’s Law** \( \frac{V}{n} = a \text{ constant} \)
- **Combined Laws** \( \frac{P V}{n T} = a \text{ const tan} t \)

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**THE IDEAL GAS EQUATION**

\[ P V = n R T \]

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**For an ideal gas**

1. the molecules of a gas have **NO volume**
2. molecules of a gas do **NOT** attract (or repel) each other.
APPLICATIONS OF THE IDEAL GAS EQUATION

1. Gas Density \( \frac{\text{weight}}{\text{volume}} \)
2. Molecular Weight \( \frac{\text{grams}}{\text{mole}} \)
3. Partial Pressures
4. Gas Volumes in Chemical Reactions

GAS DENSITY

\[
P V = \frac{\text{grams}}{MW} R T
\]

\[
P V MW = \text{grams} R T
\]

\[
\text{Density} = \frac{\text{grams}}{\text{volume}} = \frac{P MW}{R T}
\]

What happens to the density of a gas as
(a) the gas is heated in a constant-volume container;
(b) the gas is compressed at constant temperature;
(c) additional gas is added to a constant-volume container?

Answer: (a) no change,
(b) increase,
(c) increase

MOLAR MASS (MOLECULAR WEIGHT)

From \( PV = n RT \)

\[
n = \text{moles} = \frac{\text{grams}}{MW}
\]

\[
MW = \frac{gRT}{VP}
\]

Gas Mixtures and Partial Pressures

\[
\frac{P_i}{P_T} = \frac{n_i}{n_T}
\]

\[
P_i = X_i P_{\text{total}}
\]

where \( X_i \) is the mole fraction

\[
X_i = \frac{n_i}{n_{\text{Total}}}
\]

A 1.000 L sample of dry air at 25.0°C has the following composition:

\[
\begin{align*}
\text{N}_2 & \quad 0.8940 \text{ g} \\
\text{O}_2 & \quad 0.2740 \text{ g} \\
\text{Ar} & \quad 0.0152 \text{ g and} \\
\text{CO}_2 & \quad 0.0007 \text{ g}
\end{align*}
\]
What is the partial pressure of each gas and what is the Total pressure?

1st Convert grams to moles
2nd Use Ideal Gas Equation

1.000 L at 25.0°C with 0.8940 g N₂
(a) What is the partial pressure of N₂???

\[ P_N = \frac{n_N RT}{V} \]
\[ \frac{(0.8940 \text{ g}) \times \frac{1 \text{ mol}}{28 \text{ g}}}{0.0085625 \text{ mol}} \]
\[ = \frac{(0.03191 \text{ mol})(0.08206 \text{ atm} \cdot \text{L/mol} \cdot \text{K})(25 + 273)}{1.000} \]
\[ = 0.780313 \text{ atm} \]

1.000 L at 25.0°C with 0.2740 O₂
(a) partial pressure of O₂???

\[ P_O = \frac{n_O RT}{V} \]
\[ \frac{(0.2740 \text{ g}) \times \frac{1 \text{ mol}}{32 \text{ g}}}{0.0085625 \text{ mol}} \]
\[ = \frac{(0.0085625 \text{ mol})(0.08206 \text{ atm} \cdot \text{L/mol} \cdot \text{K})(25 + 273)}{1.000} \]
\[ = 0.2093748 \text{ atm} \]

The total pressure is the sum of the partial pressures

**Repeat for Argon:**  \( P = 0.0093 \)

**Repeat for CO₂:**  \( P = 0.0004 \)

\[ P_{\text{total}} = P_{\text{nitrogen}} + P_{\text{oxygen}} + P_{\text{argon}} + P_{\text{carbon dioxide}} \]
\[ P_{\text{total}} = 0.7808 + 0.2095 + 0.0093 + 0.0004 \]
\[ P_{\text{total}} = 1.000 \text{ atm} \]

**Real Gases**

**Deviations from Ideal Behavior**

**The van der Waals Equation**

\[ P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2} \]

Corrects for molecular volume
Corrects for molecular attraction
Real Gas $\rightarrow$ Ideal Gas

LOW Pressures (not much gas)
LARGE volume (lots of room between molecules)
HIGH Temperature (molecules are moving too fast to interact)

Chapter 11

**INTRA** molecular forces

**INTER** molecular forces

Chemical BONDS Formed

**BETWEEN ELEMENTS** are **INTRA** molecular forces

Chemical BONDS Formed

**BETWEEN MOLECULES (\&/or IONS)** are **INTER** molecular forces

**INTER** Molecular Forces

Much weaker than intra molecular forces.

Several types of forces:
1. Ion–dipole
2. Dipole–dipole
3. London Forces $\Rightarrow$ Instantaneous induced–dipole (dispersion forces)
4. Hydrogen “bonds.”

Section 11.2 ONLY

**THIS MATERIAL NOT COVERED IN EXAM**

Properties of Liquids
Phase Changes
Vapor Pressure
Phase Diagrams

Chapter 13 Solutions and Their Properties

- Solubility Rules – Chapter 4
- Factors that affect Solubility –
  - Pressure (Henry’s Law)
  - Temperature
- Concentration
  - Molecular weight
- Colligative Properties
  - Molecular solutions
  - Ionic solutions
TEMPERATURE EFFECTS

As temperature increases, solubility of solids generally increases

BUT Not Always

Sometimes, solubility decreases as temperature increases

PRESSURE EFFECTS

Henry’s Law: The solubility of a gas is directly proportional to the pressure of the gas over the solution.

\[ S_g = k \cdot P_g \]

CONCENTRATION EFFECTS

The amount of solute present in a given amount of solution

- IONIC Or MOLECULAR?
- STRONG ELECTROLYTES
  A substance that completely ionizes

Types of Electrolytes

1. STRONG ELECTROLYTES
   A substance that completely ionizes
   - Strong Acids [7]
   - Strong Bases [7]
   - Soluble Salts

2. WEAK ELECTROLYTES
   A substance that does NOT completely ionize
   - Weak Acids / Bases
   - Slightly Soluble Salts

Parts per Million (ppm)

Parts per million (ppm) can be expressed as 1 mg of solute per kilogram of solution

If the density of the solution is 1g/mL, then 1 ppm = 1 mg solute per liter of solution

### Table 4.1: Solubility Guidelines for Common Ionic Compounds in Water

<table>
<thead>
<tr>
<th>Soluble Ionic Compounds</th>
<th>Important Exceptions</th>
<th>Insoluble Ionic Compounds</th>
<th>Important Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing</td>
<td></td>
<td>Compounds containing</td>
<td></td>
</tr>
<tr>
<td>NO\textsubscript{3}\textsuperscript{-}</td>
<td>None</td>
<td>S\textsuperscript{2-}</td>
<td>Compounds of NH\textsubscript{4}\textsuperscript{+}, the alkali metal cations, and Ca\textsuperscript{2+}, Sr\textsuperscript{2+}, and Ba\textsuperscript{2+}</td>
</tr>
<tr>
<td>CH\textsubscript{3}HO\textsubscript{2}\textsuperscript{-}</td>
<td>None</td>
<td>CO\textsubscript{2}\textsuperscript{-}</td>
<td>Compounds of NH\textsubscript{4}\textsuperscript{+} and the alkali metal cations</td>
</tr>
<tr>
<td>Cl\textsuperscript{-}</td>
<td>Compounds of Ag\textsuperscript{+}, Hg\textsubscript{2}\textsuperscript{2+}, and Pb\textsuperscript{2+}</td>
<td>PO\textsubscript{4}\textsuperscript{3-}</td>
<td>Compounds of NH\textsubscript{4}\textsuperscript{+} and the alkali metal cations</td>
</tr>
<tr>
<td>Br\textsuperscript{-}</td>
<td>Compounds of Ag\textsuperscript{+}, Hg\textsubscript{2}\textsuperscript{2+}, and Pb\textsuperscript{2+}</td>
<td>OH\textsuperscript{-}</td>
<td>Compounds of the alkali metal cations, and Ca\textsuperscript{2+}, Sr\textsuperscript{2+}, and Ba\textsuperscript{2+}</td>
</tr>
<tr>
<td>F\textsuperscript{-}</td>
<td>Compounds of Ag\textsuperscript{+}, Hg\textsubscript{2}\textsuperscript{2+}, and Pb\textsuperscript{2+}</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
WHY STUDY CONCENTRATION?

\[ X_A = \frac{\text{Moles of } A}{\text{Total number of moles}} \]

Molarity = \( \frac{\text{Moles of solute}}{\text{Liters of SOLUTION}} \)

Molality = \( \frac{\text{Moles of solute}}{\text{Kilograms of SOLVENT}} \)

**Colligative Properties**

- **Colligative Properties**: These properties depend ONLY on the Number of solute particles in solution and NOT ON the nature of the solute particles.