CHE 1301

Basic Principles of Modern Chemistry I

Week 15

Hi! Thanks for checking out the weekly resources for Chemistry 1301! This resource covers topics typically taught by professors during the 15th week of classes.

Visit our website, https://baylor.edu/tutoring, to sign up for appointments and check out additional resources for your course! You'll find helpful links with the following titles:

- "Online Study Guide Resources" – The pace of your course may vary slightly from what's shown in this document. If you don't see the topics you're learning right now, use this link to find the weekly resources for the rest of the semester.
- "How to Participate in Group Tutoring" - See if there is a Chemistry 1301 group tutoring session being hosted this semester. These are weekly question/answer sessions taught by our master tutors!
- "View tutoring times for your course" or "Schedule a private 30-minute appointment!"

You can also give us a call at (254)710-4135, or drop in! Our hours are Monday-Thursday 9 am – 8 pm on class days.

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KEY WORDS: Study Tips, Energy Flow, Partial Pressures, Phase Diagrams

TOPIC OF THE WEEK: Study Tips/Final Review

Main idea—practice recall. Tests are made of recall, not just of understanding the material. Do make sure that everything makes sense, but **don't spend all your time rereading notes/textbook**. It's very important to practice thinking about the subject without having any resources in front of you.

1. Practice problems—see table p.3
   a. There is so much math in chemistry! Make sure you practice applying formulas.
2. Interleaved studying
   a. Instead of studying for one final, taking it, then studying for another, try setting timers and studying one subject for an hour, then studying another subject for an hour. This is helpful because it forces your mind to recall what you were working on each time you restart studying.
3. Spaced studying
   a. This is related to interleaved studying—it's the reason that interleaved studying is helpful. More short study sessions are better than fewer long study sessions because it's more recall practice.
4. Blank sheet of paper

a. Pick a topic, take a blank sheet of paper, and write down everything you know about it. The final review resource will map out the relationships between various topics. I'd suggest trying it this week and comparing with the map given next week!

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Highlight 1:

This section is very dependent on knowing definitions. The main idea is that energy is related to reactions in two ways: **they require energy to take place, and they generate energy from taking place**... the *change in internal energy* (energy contained in all reactants and products) will depend on these two values.

Calculating change in internal energy:

Method 1:

$$E_{universe} = E_{products} - E_{reactants}$$

When the *system* (this is what we call the products and reactants of whatever reaction we are discussing) loses energy, it has to go somewhere...the *surroundings* (the rest of the universe) are what gains it. Likewise, when the system gains energy, this means that it's taken the energy from the surroundings.

Method 2:

Hess's Law:

Check out this video -- https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:thermodynamics/x2eef969c74e0d802:hess-s-law/v/hess-s-law-intro

Use enthalpies of similar reactions to calculate the enthalpy of your reaction. Use these rules:

1. If you flip an equation, the energy generated becomes enthalpy released, and vice versa. In other words, change the sign.
2. If you multiply an equation by a constant, the total enthalpy of the reaction is multiplied by the same.

Method 3:

$$E_{system} = E_{products} - E_{reactants}$$

By definition, change in internal energy is the **difference** between the **final amount of energy contained by a system** and the **initial amount of energy** contained by the same. At both of these points, two types of energy are present:

1. The total amount of energy stored (potential energy) depends on the bits of energy needed to form each type of bond, because this is the same amount of energy that would be needed to break them.
2. Kinetic energy depends on how fast particles are moving. $KE = \frac{1}{2}mv^2$

a. Kinetic energy changes whenever temperature changes—warmer temp = fastermoving molecules = higher kinetic energy
b. Note: If there is a phase change, kinetic energy will change. See last week's resource!

If you are asked to calculate change in internal energy from something like this:
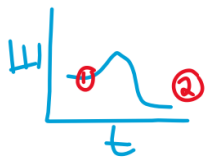
Subtract the y-value (amount of energy) at (2) from that at (1). In this case, there is less energy after the reaction than before it, which means that the amount of internal energy decreased.

==Energy released: ΔE is negative (energy in the system decreases)==

==Energy absorbed: ΔE is positive (energy in the system increases)==

This energy can be transferred in a couple of different ways—as heat or as work.

==$\Delta E = q + w$==, where q is energy transferred as heat and w is energy transferred as work

Energy flow as heat:

A lot of times, you'll see the energy lost/gained by a reaction released/inputted as heat. Here are the two terms that can describe this:

==Exothermic==—a reaction loses energy as heat (you feel the **flask heat** up, or you see a phase change in the direction **gas→liquid→solid**)

==Endothermic==—a reaction gains energy as heat (you feel the **flask cool** down, or you see a phase change in the direction **solid→liquid→gas**)

Is the melting of an ice cube endothermic or exothermic? This is a question I often got mixed up on—see Things You May Struggle With for some helpful tips.

Energy flow as work:

Work done by a gas can be calculated using ==$W = P\Delta V$==

Think about why this might make sense: if pressure increases, it is harder (more work) to change volume.

To calculate change in volume, use the formula for the volume of whatever shape the gas is contained in.

As with heat, there are two scenarios:

==Work done on a system==—energy is gained by the system. Something else does work, and now the system has the energy.

==System does work==—energy is lost by the system. The system uses some of its energy.

Summary:

System gains energy if: absorbs heat, work done on system

System loses energy if: releases heat, work done by system

Energy flow as heat cont:

Heat lost/gained by a system can be calculated in a couple of ways:

1. $q = mc\Delta t$
   a. **M = mass of the substance that has lost/gained heat.**
   b. **C = specific heat of the substance**
   c. **Δt=change in temperature**
   d. *Why would it make sense that the equation is set up this way?* If you have more mass, and a greater change in temperature, more heat will have to be added.
   e. How is this **measured in practice?**
      i. To capture the heat, scientists can use calorimeters. A bomb calorimeter makes calculations simplest, because it maintains equal pressure and volume. No energy is lost/gained due to work.
      ii. A calorimeter is a **closed system**. So, any heat lost by one part of the system (the process you're trying to determine) is gained by the other part (often, you'll see problems where this other part is water—then, we can use water's specific heat, which is 4.184 J/g degree C): $q_{lost} = q_{gained}$
2. Enthalpies of formation:
   a. $\Delta H = \Sigma n\Delta H_{products} - \Sigma n\Delta H_{reactants}$
   b. Multiply the enthalpies of formation by the number of moles of each item in the formula. Total enthalpy is the **amount of energy needed to make the products** minus the **energy already in the system due to the reactants**.

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Highlight 2: Partial Pressures

In your average random gas, there are a lot of different types of molecules that make up the one *mixture*. And all of them contribute to the pressure, since all of their molecules are hitting the "walls of the container" (see Week 13 to review Properties of Gases). But their pressures can be calculated individually—P_A means, **how much is Gas A hitting the walls of the container?** The important thing to remember is that $P_{total}=P_A+P_B+P_C...$

Also, each partial pressure is proportional to its mole fraction of the whole mixture. $P_A P_{tot}=n_A n_{total}$

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Highlight 3: Periodic Trends

Phase diagrams plot **states of matter** at various **temperatures** and **pressures**. Areas of the graph are blocked off, because states of matter don't just depend on temperature, or just on pressure— the two together decide the state of matter. Here is the phase diagram of water, as an example (from Chemistry – An Atoms-First Approach (3rd Ed.), Zumdahl, Zumdahl, & Decoste.)
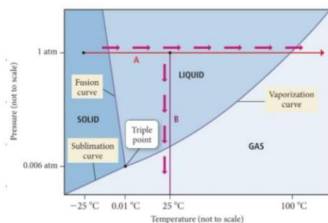
You may be asked to identify different parts of the plot. Here are some labels to know:

Triple point – the temperature and pressure at which the substance exists as all 3 states of matter, simultaneously

Fusion curve—the line that blocks off [pressures/temperatures that make the substance a solid] from [pressures/temperatures that make the substance a liquid]

Vaporization and sublimation curves—the same sorts of divisions, but between liquids/gases and solids/gases

Critical point—point above which phase division no longer exists…if a substance gets above the **critical temperature**, it will no longer be able to turn back into a liquid. The critical pressure is the pressure that it takes to turn a gas into a liquid at the *critical temperature.*

A key thing to keep in mind about the **rate of change of vaporization**:

Vapor pressure: the partial pressure (see Highlight 2) of a given molecule, above its liquid form, once it reaches equilibrium:

A certain amount of **any given liquid** substance **evaporates** into **gaseous** form. When this happens, some of the **gaseous** substance **recondenses** and becomes **liquid**. As more and more of the liquid becomes gas, the **rate of condensation increases** to try to compensate. When the **partial pressure of the substance** equals the **pressure applied by the atmosphere**, the rate of condensation stops increasing. This point is identical to boiling point.

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Check Your Learning

1. Given the following enthalpies of formation, calculate the total enthalpy of the reaction: a. Reactant A: 246 kJ, Product -400 kJ, Reactant A 558 kJ
2. Hess's Law: https://general.chemistrysteps.com/hesss-law-practice-problems/
3. If the flask heats up, is the reaction exothermic or endothermic? Does it gain or lose energy?
4. How would you find the triple point on a phase change diagram?

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Things You May Struggle With

1. Endothermic and exothermic reactions: If an ice cube melts, has the system containing the ice cube undergone an endothermic or exothermic process? The answer is endothermic, but this seems strange. Aren't endothermic reactions "cold" reactions? They can be, but it depends on the reference frame! If you feel a flask containing an endothermic reaction, it will be cold. But remember—the flask is part of the surroundings. If the flask is cold, energy has gone into the system—energy has gone into the reaction. When figuring out whether a process is exothermic or endothermic, think about whether energy has been released or gained.
2. Quick note—phase changes are physical processes, not reactions. However, all of these processes can be classified as exothermic/endothermic.

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That's it for this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to Check Your Learning are below.

1. 404 kJ/mol
2. See link
3. Exothermic, loses/releases to surroundings
4. Where all 3 lines meet