1. If a 40.1 g piece of iron at 652 °C is dropped into a sample of 328 g of water at 32.4 °C, what will the final temperature be after thermal equilibrium is established? Assume that no heat is lost during the process.

Answer: 40. °C

Analysis of the situation: What happened: 328 g of water went from 32.4 °C to some (as yet unknown) final temperature ($T_{f,\,water}$) as 40.1 g of iron (Fe) went from 652 °C to some (as yet unknown) final temperature ($T_{f,\,Fe}$).

Conclusions from the above:
   1) Since the iron was initially hotter than the water, some heat energy q must have transferred from the iron to the water during the described process.
   2) The value of q that flowed into the water must equal the value of q that flowed out from the iron (since there was only one "amount" of heat that flowed). In mathematical terms, this can be stated as: $q_{water} = -q_{Fe}$ (or $q_{Fe} = -q_{water}$; either is fine.)
   3) Since the two samples end up at thermal equilibrium, the two final temperatures noted above must be the SAME temperature. That is, $T_{f,\,water} = T_{f,\,Au}$, and so I will refer to these both as simply T_f from here on out in this problem.

Strategy:
   1) Use the knowledge (equation) that the amount of heat transferred from (or to) a sample is equal to the product of its specific heat, mass, and change in temperature. Write this product for both q_{water} and q_{Fe} (obtaining the specific heat capacity of iron from a table in your text)
   2) If you then set q_{water} equal to the opposite of q_{Fe} (see (2) above in conclusions), you will see that there is only one unknown left to solve for: T_f.
   3) Use algebra to solve for the value of T_f.

Execution of Strategy:

$$q_{water} = -q_{Fe}$$

$$C_{s,water} \times m_{water} \times \Delta T_{water} = -\left(C_{s,Fe} \times m_{Fe} \times \Delta T_{Fe}\right)$$

$$\Rightarrow \quad \frac{4.184 \text{ J}}{g \cdot {}^{\circ}C} \times 328 \; g \times \left(T_f - 32.4 \; {}^{\circ}C\right) = -\left[\frac{0.45 \text{ J}}{g \cdot {}^{\circ}C} \times 40.1 \; g \times \left(T_f - 652\right)\right]$$

$$\Rightarrow \quad \frac{137\underline{2}.4 \text{ J}}{{}^{\circ}C} \times \left(T_f - 32.4 \; {}^{\circ}C\right) = -\frac{1\underline{8}.05 \text{ J}}{{}^{\circ}C} \times \left(T_f - 652\right)$$

Dividing through by $1\underline{8}.05$ J/°C yields (you could do the algebra other ways as well [i.e., distribute on both sides]):

$$\frac{\left(137\underline{2}.4 \text{ J/}{}^{\circ}C\right)}{\left(1\underline{8}.05 \text{ J/}{}^{\circ}C\right)}\left(T_f - 32.4 \; {}^{\circ}C\right) = -\left(T_f - 652 \; {}^{\circ}C\right)$$

$$\Rightarrow \; 7\underline{6}.0(T_f - 32.4 \; {}^{\circ}C) = -T_f + 652 \; {}^{\circ}C$$

$$\Rightarrow \; 7\underline{6}.0 T_f - 2\underline{4}63 \; {}^{\circ}C = -T_f + 652 \; {}^{\circ}C$$

$$\Rightarrow \; 7\underline{7}.0 T_f = 652 \; {}^{\circ}C + 2\underline{4}63 \; {}^{\circ}C = 3\underline{1}15 \; {}^{\circ}C$$

$$\Rightarrow T_f = \frac{3115\ °C}{77.0} = 40.45\ °C = \boxed{40.\ °C}$$

> NOTE: At first it may seem like the value of T_f determined is way too small. It is only ~8 °C above the initial T of the water! But this IS correct. The reason that the temperature change for the iron is so much bigger (in absolute value) than the ΔT for the water is due to two factors: a) The specific heat of water is more than NINE TIMES larger than that of iron (4.184 vs 0.45 J/g°C). That means it takes much less energy (per gram) to raise iron's temperature, which also means you get back much less energy (per gram) for each degree of cooling. b) The mass of the water in this case was a little more than EIGHT TIMES that of the iron (328 g vs. 40.1 g). Taken together, these two factors made the temperature change of the iron more than 70 times greater (9 x 8 = 72) than the temperature change of the water!

2. How much heat (in kJ) is evolved or absorbed in the reaction of 233.0 g of carbon with enough CaO to produce calcium carbide? (b) Is the process exothermic or endothermic?

$$CaO(s) + 3\ C(s) \rightarrow CaC_2(s) + CO(g); \quad \Delta H° = 464.8\ kJ$$
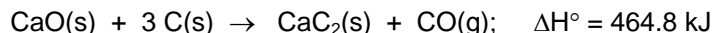
(a) **Answer**: 3006 kJ

Strategy: 1) Determine the number of moles of C in 233.0 g of it (by using the molar mass of C).

2) Recognize that the "464.8 kJ" here means that "464.8 kJ is absorbed when *three moles* of C(s) reacts with 1 mol of CaO(s) to produce....." and use that ratio to figure out how much energy is absorbed when *x* moles of C(s) are reacted (where *x* is the number determined in (1)).

Execution of Strategy:
1) The molar mass of C is 12.01 g/mol. So 233.0 g is roughly 20 moles (233 is about 20 times bigger than 12). More precisely, it is 233.0g /(12.01 g/mol) = 19.40 mol or:

$$233.0\ g \times \frac{1\ mol\ C}{12.01\ g} = 19.400\ mol\ C$$

2)
$$19.400\ mol\ C \times \frac{464.8\ kJ \quad (of\ heat\ absorbed\ )}{3\ mol\ C \quad (reacted\ )} = 3005.7 = \boxed{\textbf{3006 kJ}\ absorbed}$$

(b) The process is **endothermic.** A positive value of ΔH_{sys} indicates that H increases for the system, and thus energy flows *in* as the process occurs. That describes **endo**thermic.

3. 85.8 kJ of energy is evolved (i.e., released) at constant pressure when 3.56 g of P_4 is burned according to: $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$ What is the ΔH for the (thermo)chemical equation?

Answer: -2.99 x 10³ kJ

Reasoning:

Since the coefficient of P_4 is "1" in the balanced equation, you need to find the amount of energy released when ONE MOLE of P_4 is burned to get the <u>magnitude</u> of the ΔH for the (thermo)chemical equation. How many moles is 3.56 g of P_4? Molar mass of P_4 = 4(30.97 g/mol) = 123.9 g/mol P_4. So:

$$\frac{3.5\underline{6}\text{ g}}{123.\underline{9}\text{ g/mol}} = 0.028\underline{7}3\text{ mol }P_4 \text{ was combusted}$$

Thus the energy released PER MOLE is just (remember the "How many numbers do you see in the fraction on the left below? Just one....you just haven't done the division yet!" idea):

$$\frac{85.8\text{ kJ released}}{0.028\underline{7}3\text{ mol }P_4\text{ burned}} = \frac{\left(85.8\middle/0.028\underline{7}3\right)\text{kJ released}}{\text{mol }P_4\text{ burned}} = \frac{29\underline{8}6\text{ kJ released}}{\text{mol }P_4\text{ burned}} = 2.99 \times 10^3 \text{ kJ (released per mol }P_4\text{)}$$

Since energy is *released*, the <u>sign</u> of ΔH is *negative*. Thus, $\Delta H = \boxed{\textbf{-2.99 x 10}^\textbf{3}\textbf{ kJ}}$.

4. Instant cold packs contain solid NH_4NO_3 and a pouch of water. When the pack is squeezed, the pouch breaks and the solid dissolves, lowering the temperature because of the endothermic process:

$$NH_4NO_3(s) \rightarrow NH_4NO_3(aq); \quad \Delta H = +25.7\text{ kJ}$$
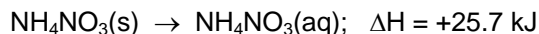
What is the final temperature in a squeezed cold pack that contains 50.0 g of NH_4NO_3 dissolved in 125 mL of water? Assume a specific heat of 4.18 J/(g·°C) for the solution, an initial temperature of 25.0 °C, and no heat transfer between the cold pack and the environment. (Recall, d_{water} ~ 1.0 g/mL)

<u>**Answer**</u>: $\boxed{\text{-5.73 °C (or 3.1 °C; see comment below*)}}$

<u>**Strategy**</u>: Recognize that this is a "heat flow" kind of problem, which means:

$$q_{sys} = -q_{surr}$$

Since a *process* is involved, let the system be the atoms/molecules/ions that are involved in the process. Here that would be the NH_4NO_3 (and the few water molecules that solvate the ions, once dissolved). Thus q_{sys} becomes $q_{process} = \Delta H_{sys}$. The surroundings is thus the (bulk) water (or the whole "cold pack"; see comment*). Thus the above equation simplifies to:

$$\Delta H_{sys} = -q_{water} \quad \text{(or } -q_{cold\,pack}\text{; see comment*)} \quad \text{which becomes}$$

$$\Delta H_{sys} = -(C_{s.water} \times m_{water} \times \Delta T_{water}) \quad \text{(where } \Delta T_{water} = T_{final} - T_{initial} \text{ for the water)}$$

Knowing the mass of NH_4NO_3 that "reacts" (strictly speaking this is actually a *physical* change, but we can view the solid as a "reactant" here) and knowing the enthalpy change for when *one MOLE* of it "reacts" (i.e., understanding the meaning of a ΔH that comes after a chemical equation) allows you to figure out ΔH_{sys} (i.e., how much energy was absorbed by the system as <u>*this specific amount*</u> of dissolution (50 g) takes place). m_{water} can be obtained from V_{water} and the density of water. s_{water} is given (see comment*) and $T_{initial}$ is given. Thus the only unknown is T_{final}.

<u>**Execution of Strategy**</u>:

    1) 50.0 g/(80.0 g/mol) = 0.625 mol NH_4NO_3

2) $(0.625 \text{ mol } NH_4NO_3)(25.7 \text{ kJ/mol}) = 16.\underline{0}6 \text{ kJ} = 16.\underline{0}6 \times 10^3 \text{ J absorbed} = \Delta H_{sys}$

3) $\Delta H_{sys} = -q_{water}$

$$16.\underline{0}6 \times 10^3 \text{ J} = -\left[\frac{4.18 \text{ J}}{g \cdot °C} \times \left(125 \text{ mL} \times \frac{1.00 \text{ g}}{mL} \right) \times \left(T_f - 25.0 °C \right) \right]$$

$$\Rightarrow \quad -\frac{16.\underline{0}6 \times 10^3 \text{ J}}{\left(4.18 \enspace J\!\!\Big/\!_{g \cdot °C} \right)(125 \text{ g})} = T_f - 25.0 °C \quad \Rightarrow \quad -30.\underline{7}3 °C = T_f - 25.0 °C$$

$$\Rightarrow T_f = -30.\underline{7}3 + 25.0 °C = \boxed{\textbf{-5.73 °C}} \textbf{ (SEE NOTE BELOW*)}$$

*NOTE: If you assume the specific heat capacity of the *solution* is 4.18 J/(g·°C) [in effect, if you assume that the specific heat capacity of the dissolved salt is the same as that of water, which it clearly isn't] and use a *solution* mass of 175 g (i.e., including the mass of the NH_4^+ and NO_3^- ions), the answer obtained is about **3.1 °C** for the final temperature. The actual "answer" is likely some value in between--I think it would be closer to -5.7 than +3.1 because the specific heat capacity of the ions is likely to much less than that of water, whose C_s value is quite high compared to just about any other substance. In lab, we IGNORED the mass of the dissolved solutes when doing this sort of calculation which is why I did the current calculation that way in this key.

5. **8.45.** Assume that a particular reaction produces 244 kJ of heat and that 35 kJ of *PV* (expansion/contraction) work is done on the system. What are the values of ΔE and ΔH for the system? For the surroundings?

Answers: $\boxed{\Delta E_{sys} = -244 + 35 = \textbf{-209 kJ}; \ \Delta H_{sys} = \textbf{-244 kJ}; \ \Delta E_{surr} = \textbf{+209 kJ}; \ \Delta H_{surr} = \textbf{+244 kJ}}$

Reasoning: The energy of a system can only be changed by two things: addition (or loss) of heat (energy) or by work being done on (or by) the system. Another way to say this is that E_{sys} can go down if heat energy is lost (to the surroundings) and/or if work is done *by* the system (on the surroundings); E_{sys} can go up if heat energy is gained (from the surroundings) and/or if work is done *on* the system (by the surroundings). So if 244 kJ of heat is "evolved" by the system, this means that 244 kJ of heat energy is lost, and so E should go down by 244 kJ as a result of this change. If 35 kJ of work is gained by the system, E should go up by 35 kJ as a result of this change. If both changes occur, the result is the sum of the two: -244 kJ + 35 kJ = -209; a net loss of 209 kJ (by the system to the surroundings) results. So $\Delta E_{sys} - 209$ kJ and $\Delta E_{surr} = +209$ kJ. Since $\Delta H = q_p$, ΔH_{sys} is simply -244 kJ here and ΔH_{surr} is the opposite of this: +244 kJ.

Using an equation rather than verbal reasoning: $\Delta E_{sys} = q + w = -244 \text{ kJ} + 35 \text{ kJ} = -209$ kJ (the other results follow)